ORAU

# PHYSICS OF TECHNOLOGY

COORDINATED BY AMERICAN INSTITUTE OF PHYSICS

# THE FLUORESCENT LAMP

Atomic Physics and Atomic Spectrum.

# THE FLUORESCENT LAMP

A Module on Atomic Physics and Atomic Spectrum

ORAU

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**The Fluorescent Lamp**



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# THE FLUORESCENT LAMP

## PREFACE TO THE STUDENT

This module is a study of the fluorescent lamp, but more especially of the light given off by such a lamp. It includes experiments with other kinds of sources of light, especially those involving gases. To explain the experiments, the structure of matter is explored. The light emitted (or absorbed) by a gas gives a great deal of information about the gas, and this is examined in some detail. Finally, the various processes taking place in the fluorescent lamp are explored.

The module is written with the idea that, if you wish, you can proceed through it with minimum help from your instructor. The laboratory instructions, questions, and problems are scattered through the text at appropriate points.

Some questions are meant to stimulate *further* thought. They cannot be answered merely by quoting the text. You should be able to answer them, however, by a combination of an *understanding* of the text material, some *common sense,* and perhaps some *previous knowledge* of science. The ability to give clear answers to such questions is an important goal of a study of any subject. We hope you will enjoy working with this module.

## PREREQUISITES

Before you start on this module, you should understand a few ideas about energy, work, and electric charge. Specifically, you should be familiar with:

1. Kinetic energy and how to find it ($\frac{1}{2}mv^2$).
2. How to find changes in gravitational potential energy ($Wh$).
3. Conservation of energy.
4. How to find the work done when a force moves an object in a straight line ($Fd$).
5. The relationships between energy and work.
6. Electric current.
7. Electrical potential difference (voltage).
8. The attraction and repulsion of electric charges.
9. Potential energy of an electric charge in an electric field.

If you are not already familiar with these ideas, you may learn about them from other Physics of Technology modules, such as *The Cathode Ray Tube,* or from other sources. The following prerequisite test will tell you where you stand. If you can answer all the questions, you are certainly ready to start on this module. If you have trouble with some of them, get some help from your teacher or another student.

## PREREQUISITE TEST

1. Which of the following possesses kinetic energy?
   a. A bird flying through the air.
   b. A tennis ball moving with a speed of 50 ft/s.
   c. A stretched spring (motionless).
   d. The moon.
   e. The flywheel in an automobile engine (engine running).

2. If a car which is moving at a constant speed of 15 mph accelerates to 30 mph, its kinetic energy
   a. does not change.
   b. becomes twice as large as before.
   c. becomes four times as large as before.
   d. becomes half as large as before.

3. A Toyota and a Buick which has twice the mass of the Toyota are both traveling at a speed of 20 mph. The kinetic energy of the Buick is
   a. the same as that of the Toyota.
   b. twice as large as that of the Toyota.
   c. four times as large as that of the Toyota.
   d. half as large as that of the Toyota.

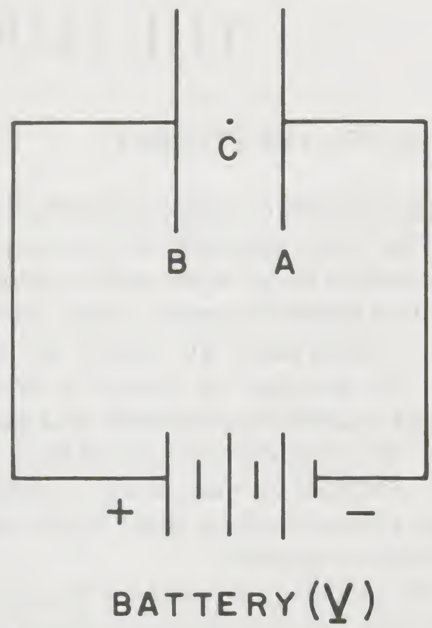
4. Which of the following possess potential energy?
   a. A bird flying through the air.
   b. A stretched spring (motionless).
   c. A storage battery.
   d. A pole vaulter at the top of his vault.

5. The amount of work necessary to lift a box which weighs $W$ to a height $h$ is
   a. $Wh$.
   b. $W/h$.
   c. $Whg$.
   d. none of the above.

6. If a box is raised to a different level, the work which was done on the box is changed into
   a. heat.
   b. kinetic energy of the box.
   c. potential energy of the box.

7. If a box is allowed to fall, which of the following statements are true?
   a. As the box falls, its potential energy is changed into kinetic energy.
   b. Just before the box hits the ground, its kinetic energy is greatest.
   c. While the box is in the air, the sum of the kinetic and potential energies of the box is constant.

Questions 8 - 12 refer to the circuit shown schematically. A and B indicate two flat sheets of metal which are parallel to each other. They are called *plates*. The combination of A and B is called a *parallel plate capacitor*.




8. The plate labeled A is
   a. negatively charged.
   b. positively charged.
   c. uncharged.

9. The potential difference between the capacitor plates is
   a. $V$.
   b. $V + IR$.
   c. $V - IR$.
   d. none of the above.

10. A negatively charged article placed at point C will
    a. move toward plate A.
    b. move toward plate B.
    c. remain at point C.

11. A negatively charged particle placed between the plates of the capacitor will possess its greatest potential energy when it is near
    a. plate A.
    b. plate B.
    c. plate C.

12. The amount of work done on a negatively charged particle which is placed near plate A and allowed to accelerate to

plate B depends on

a. the distance from A to B.
b. the current in the circuit.
c. the potential difference, $v$ (voltage).

13. Indicate whether each of the following statements is true or false.
    a. Positive charges repel negative charges.
    b. The current in a circuit depends only on the amount of charge which flows through the circuit.
    c. The current in a circuit depends both on the amount of charge which moves through the circuit and on how fast the charge moves.

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## GOALS

When you have finished this module, you should be able to:

1. Analyze light from fluorescent lamps and other sources.

2. Describe light in a precise way.

3. Describe why light from a fluorescent lamp is similar to, and yet differs from, light from other sources.

4. Describe an atom in simple terms.

5. Describe the role played by energy in the interactions between light and atoms.

6. Make use of the light given off by a gas to learn about the structure of the atoms in the gas.

7. Describe how some materials can absorb light of one color and give off light of another color.

8. Relate atomic structure to the processes which contribute to light emission from a fluorescent lamp.

# SECTION A

## Some Basic Ideas and Tools for the Study of Light

### Introduction

This module is about the principles of *atomic physics.* You will learn about atomic physics by observing the operation of various kinds of lamps, especially fluorescent lamps. You will start with naked-eye observations of the light given off by various lamps. You will then move to a more detailed analysis, using special instruments. Observations you make will then be related to the behavior of atoms.

### A First Look at Lamps and Spectra

Figure 1 is an illustration of a panel on which are mounted two common types of commercial lamps. They have been chosen to look as much alike as possible.*

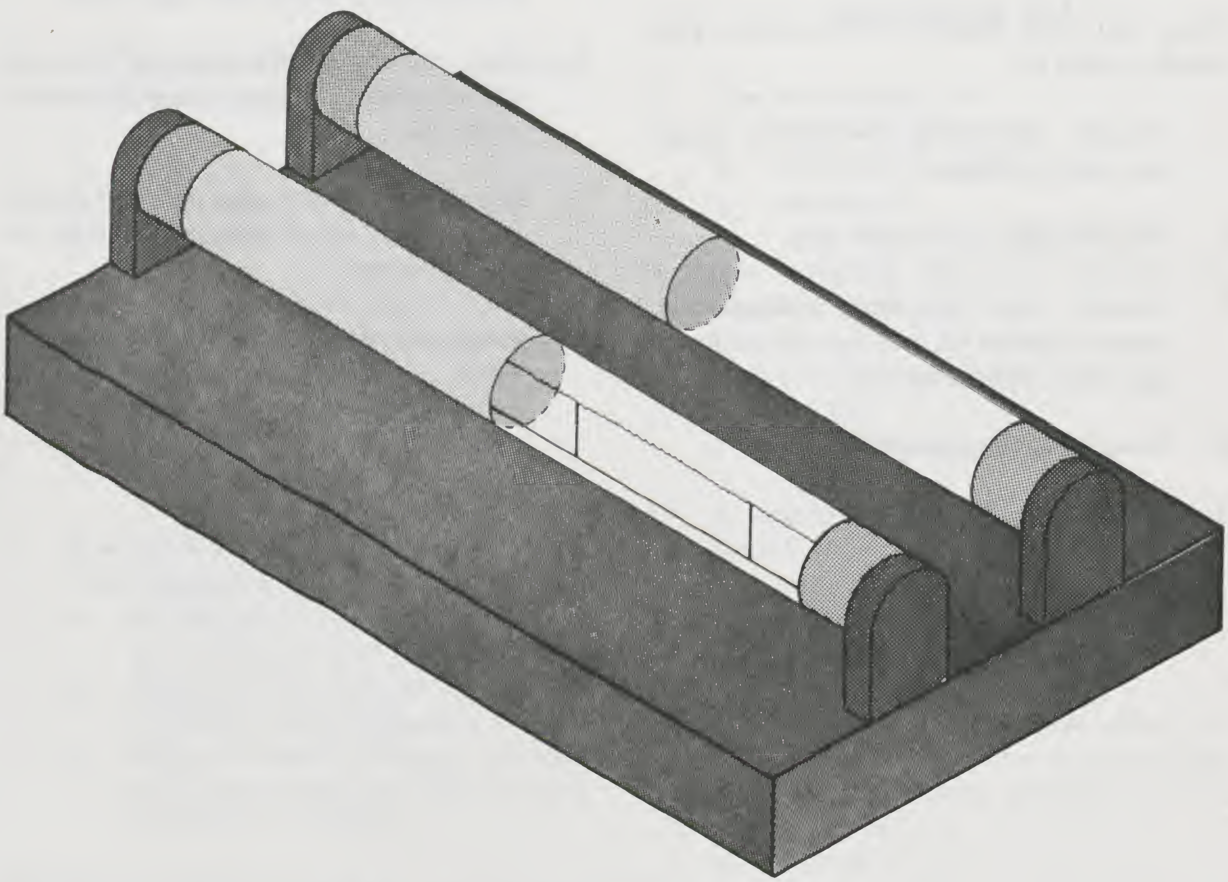

Figure 1. Two lamps, arranged for comparison studies.

*The apparatus which is provided for you may be a bit different from this. Instead of single tubes which are half frosted and half clear, you may have pairs of tubes which are identical except that one tube of the pair is completely frosted and the other is completely clear. If this is the case, adjust the following directions accordingly.

**EXPERIMENT A-1. A First Look at Lamps and Spectra**

A. Slide the covers so that the clear end of each lamp is covered. Turn both lamps on and compare the frosted ends with your naked eye. For one of the lamps you will have to turn on the switch first, then hold down the button for a few seconds to light it. Is the frosting on the inside or on the outside surface of the glass? Are they the same color? Which is brighter? Using your hand, decide which one gives off more heat into the air above it. Which one feels hotter when you touch it? Now slide the covers so that the frosted ends of the two lamps are covered and look at the clear ends. Again compare the two lamps carefully. Is the light from the lamps the same color? Which is brighter? Examine the internal structure of each lamp. Can you identify the source of the light for each lamp?

One lamp has a metal wire, or *filament,* extending from one end to the other. This filament is clearly the source of the light which this lamp emits. The light is given off when the filament is heated to a high temperature by electrical energy. A lamp which uses a heated filament as the source of light is called an *incandescent lamp.*

The lamp which requires two switches appears to be an empty glass tube. When it is on, the light seems to come from empty space within the tube. Actually, the tube is not empty, but contains materials in the form of a gas. This lamp is a *fluorescent lamp.*

Although with the unaided eye it is possible to see some difference between the uncoated ends of the two lamps, there is a tool, called a *diffraction grating,* that will yield more information. The diffraction grating "analyzes" light in the same way you analyze a complex problem, by breaking it down into simpler parts. It is not necessary to understand how the grating works in order to use it.

B. Turn on the incandescent lamp. Hold the diffraction grating, which is in the cardboard frame, close to your eye and look through it at the clear end of the incandescent lamp. Be sure to hold the grating so that the arrows on its sides are parallel to the length of the lamp (Figure 2).

Use crayons to record your observations, indicating both the color and width of each part. Now slide the cover on the lamp so that you can observe the frosted end of the incandescent lamp. Again record your observations carefully. How do the two observations compare?

The diffraction grating breaks light into parts, recognizable as colors. The spreading out of the colors is called the *spectrum* of the light. The first spectrum ever seen by most people is a rainbow, which is produced by sunlight striking raindrops. The rainbow spectrum is one in which one color blends smoothly into the next. The order, from red to orange to yellow to green to blue, is always the same. It is called a *continuous spectrum.* The light from the incandescent lamp forms a continuous spectrum.

C. Turn on the fluorescent lamp and observe the frosted end through the diffraction grating. Again keep the arrows on the grating parallel to the axis of the lamp. Record what you see. Compare this spectrum with that of the incandescent lamp. Is it a continuous spectrum?

By now you may have noticed a certain symmetry in what you see through the grating. That is, what you see on one side of the light source is reproduced on the other side. As you look to either side of the source, you always see the same color sequence. You may also have observed that the colors seen on one side are faintly repeated farther away on that same side. When making observations, you should concentrate your attention on only one of the several spectra you see.

D. With the fluorescent lamp on, look at its clear end through the diffraction grating. Again, record your results, both in writing and with crayons. How does this spectrum compare with those previously observed?

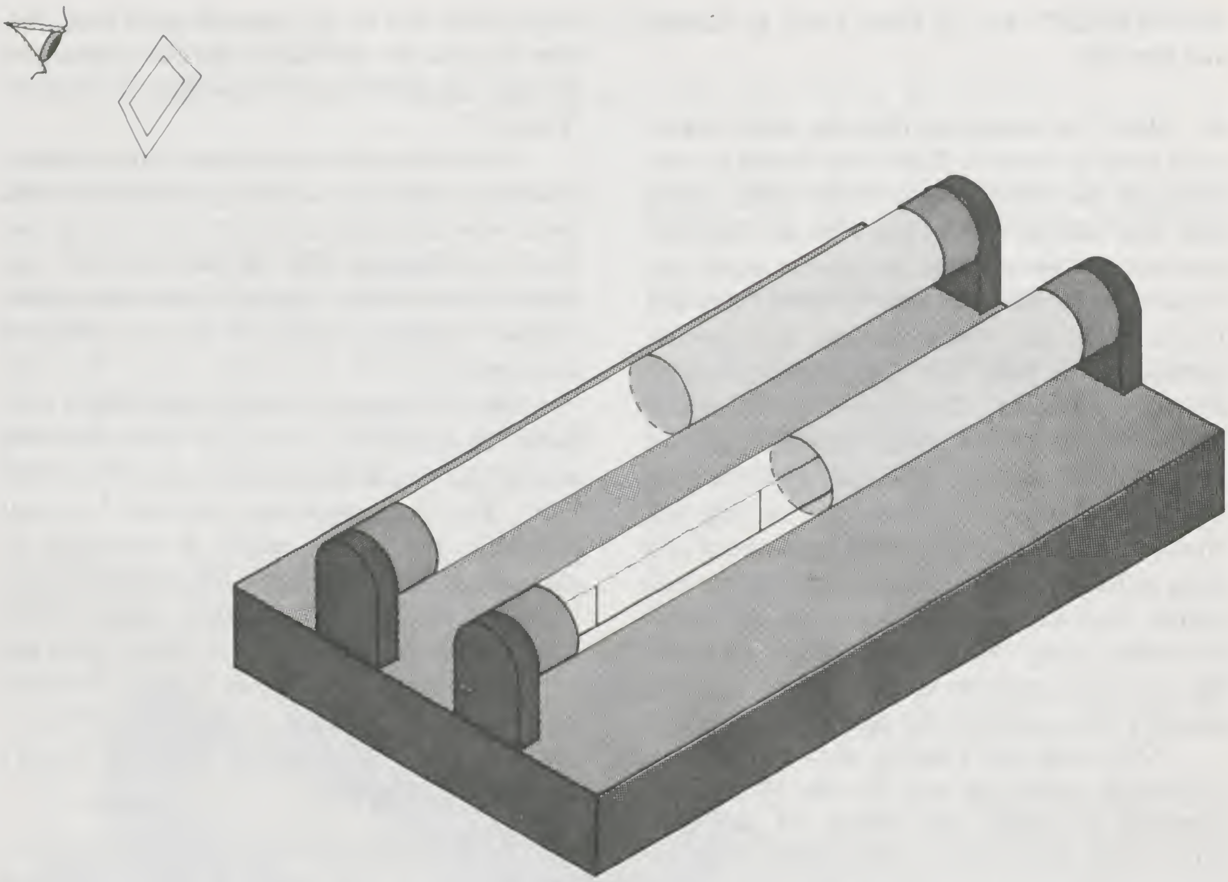

Figure 2. Viewing a lamp with a diffraction grating.

Instead of a continuous smear of color, you observed separate colored images of the clear end of the fluorescent lamp. This group of images is an example of what is called a *line spectrum.* Each colored image is called a *line.*

**Question 1.** Suppose an incandescent lamp is covered with a piece of blue cellophane. What would you expect to see through the diffraction grating? How would it compare with the blue light of the clear end of the fluorescent lamp?

**Question 2.** Look at your record of all the spectra you have studied. Is there a particular order in which the colors appear? Which color appears to be closest to the source? Which color is farthest away?

## EXPERIMENT A-2. The Spectrum of a Flame

You have used electrical energy to create light in both the incandescent and the fluorescent lamps. Light is also given off by flames. So let's use the diffraction grating to analyze the light from a flame.

Place a wire screen above a bunsen burner and light the burner in a darkened room. Adjust the air flow so that the flame gives off as little light as possible. The flame should be almost transparent and pale blue in color, with little or no yellow color. Use tweezers to place several pieces of rock salt on the screen. You should see a bright yellow glow, which is caused by the presence of sodium gas in the flame.

The sodium is found in the salt. This yellow light is a means of identifying the presence of sodium in the flame. Other materials would produce flames of different colors.

Use the diffraction grating (Figure 3) to observe the light from the flame. Hold the grating so that the arrows are vertical and record what you see. How does this spectrum differ from those previously observed? How is it similar?

---

**Question 3.** Would you describe the sodium flame spectrum as a "line spectrum" or as a "continuous spectrum"? Why?

You might wonder if the yellow seen in the line spectrum of the fluorescent lamp is the same as the yellow seen in the sodium spectrum. Perhaps your answer depends on how well you can distinguish between different shades of yellow. A color you call yellow may appear to be orange to someone else. Since the color you see depends on your own vision, color is not a very precise way to describe a spectrum. To a scientist, precision usually implies accurate measurement. When you can measure something, you can give a more precise description of it.

So, how can we measure the "color" of light? To do this we need to discuss the nature of light.

## LIGHT AS A KIND OF WAVE

Many experiments have led to the view that light is *wavelike* in nature. (The fact that the diffraction grating works as it does is one indication of this.) Light has many of the same properties as other kinds of waves.

You have probably been fascinated by ripples (waves) on the surface of a lake or a pond. If a stone is thrown into a quiet pond, the resulting disturbance (wave) spreads across the pond's surface. It moves outward at a constant speed (Figure 4).

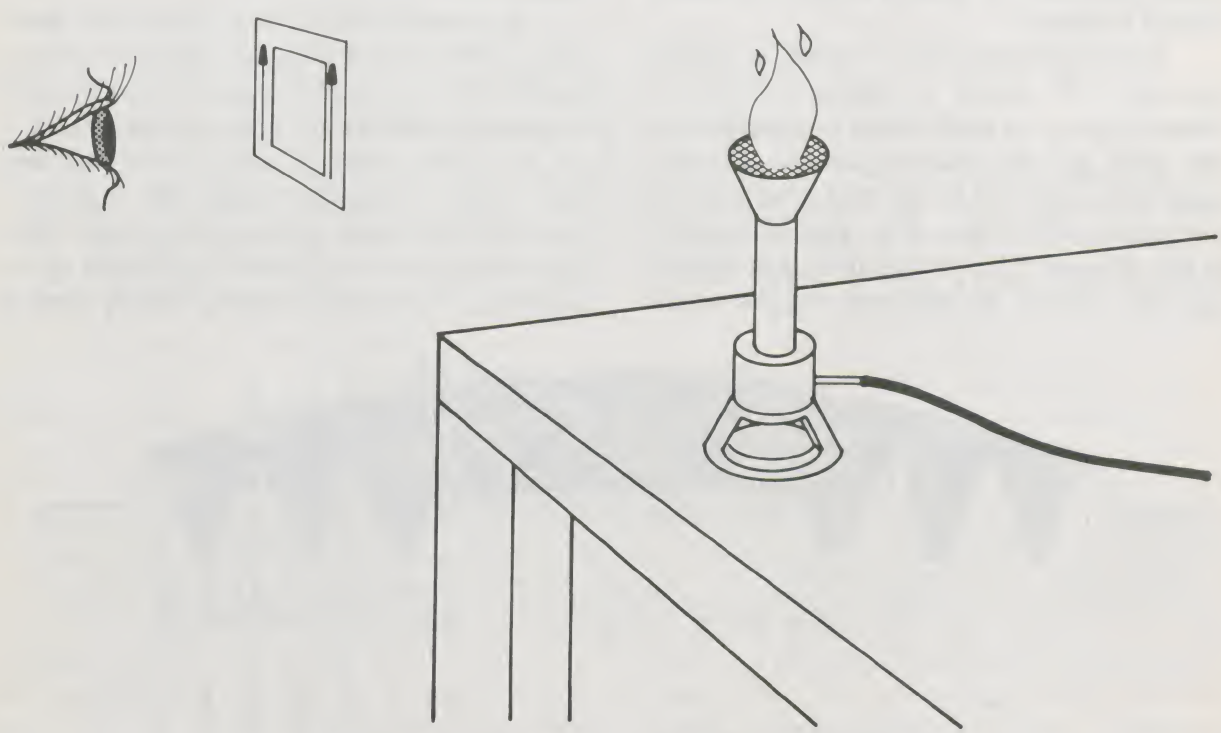

Figure 3. Viewing sodium vapor in a flame.

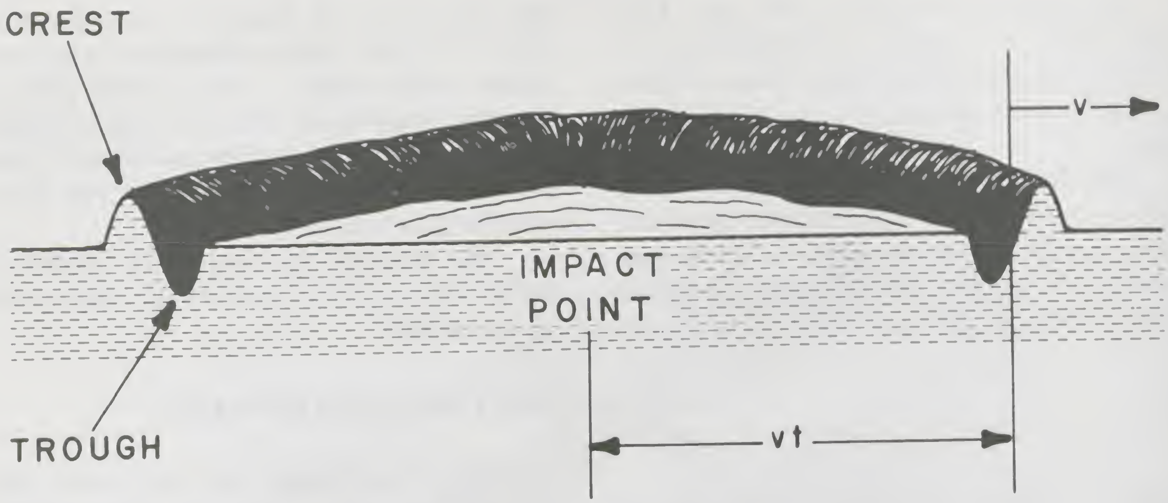


Figure 4. A water wave spreading outward.

Light, too, has been shown to move with a constant speed of about $3 \times 10^8$ m/s (186,000 mi/s) in empty space, and more slowly in matter.

Each water wave has a *crest* (the point where the water is raised highest above the normal position) and a *trough* (the point where the water is pushed farthest below its normal position).

Dropping stones into a pond at regular intervals will create a regular pattern of circular waves. A single wave is composed of one crest and the adjacent trough. The distance from one crest to the next crest is called the *wavelength* of the wave. This is the same as the distance from trough to trough (Figure 5). The number of crests or troughs which pass any given point in one second is called the frequency of the wave. We will use the Greek letter *lambda* ($\lambda$) to represent wave length, and the letter ($f$) for frequency.

The frequency, the wavelength, and the speed are related to one another in a simple fashion, which we shall now derive (see Figure 6).

In a time $\tau$, the crest of a particular wave moves from A to B, which is a distance of one wavelength, $\lambda$. Since distance traveled is equal to speed multiplied by time, we can write $\lambda = v\tau$. A person sitting at point B and watching the waves travel past would see a certain number, $f$, of crests pass in one second. How is $f$ related to $\tau$? To answer this, think of an example: If a person's heart beats 60 times a

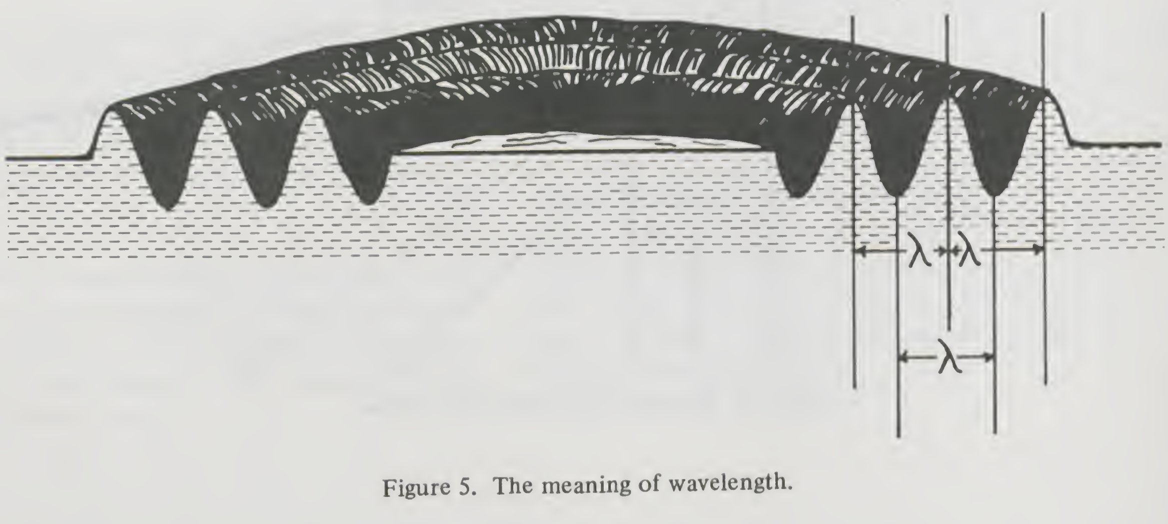


Figure 5. The meaning of wavelength.

minute, the time between beats is 1/60 minute. If the time between beats were 0.01 minute, then the pulse rate would be 100 per minute. Likewise, if the person sees $f$ crests pass in a second, the time between crests is $1/f$ seconds. This is just the time for a complete wave to travel a distance of one wavelength. So $\tau = 1/f$ and

$$\lambda = v(1/f)$$

or

$$v = f\lambda \qquad (1)$$

That is, the speed at which a wave travels is equal to the product of frequency and wavelength. This is an important equation which is valid for all kinds of waves. Since wavelength is a length, the unit for $\lambda$ is the meter. Because light has a very short wavelength, $\lambda$ for light is usually measured in nanometers (nm). (1 nanometer = $10^{-9}$ meter.) The frequency $f$ is usually measured in *hertz* (Hz). A frequency of one hertz means one wave per second.

To see how this equation can be used, let's apply it to sound waves. The speed $v$ at which sound waves travel in air is about 340 m/s. To find the wavelength of a sound wave of frequency 20 kHz (20 *kilohertz* or 20,000 Hz), we proceed as follows. Since wavelength ($\lambda$) multiplied by frequency ($f$) gives speed ($v$), we can solve for wavelength:

$$\begin{aligned}\lambda = v/f &= (340\ \text{m/s})/20\ \text{kHz} \\ &= (340\ \text{m/s})/(20 \times 10^3/\text{s}) \\ &= 17 \times 10^{-3}\ \text{m} = 1.7\ \text{cm}\end{aligned}$$

A frequency of 20 kHz is a slightly higher frequency than many people can hear.* In other words, a 1.7 cm wavelength is the shortest most people can hear.

The shortest wavelength most people can see is about $4 \times 10^{-7}$ m (400 nm). The frequency of such radiation is found as in the above example. Solving the equation for frequency:

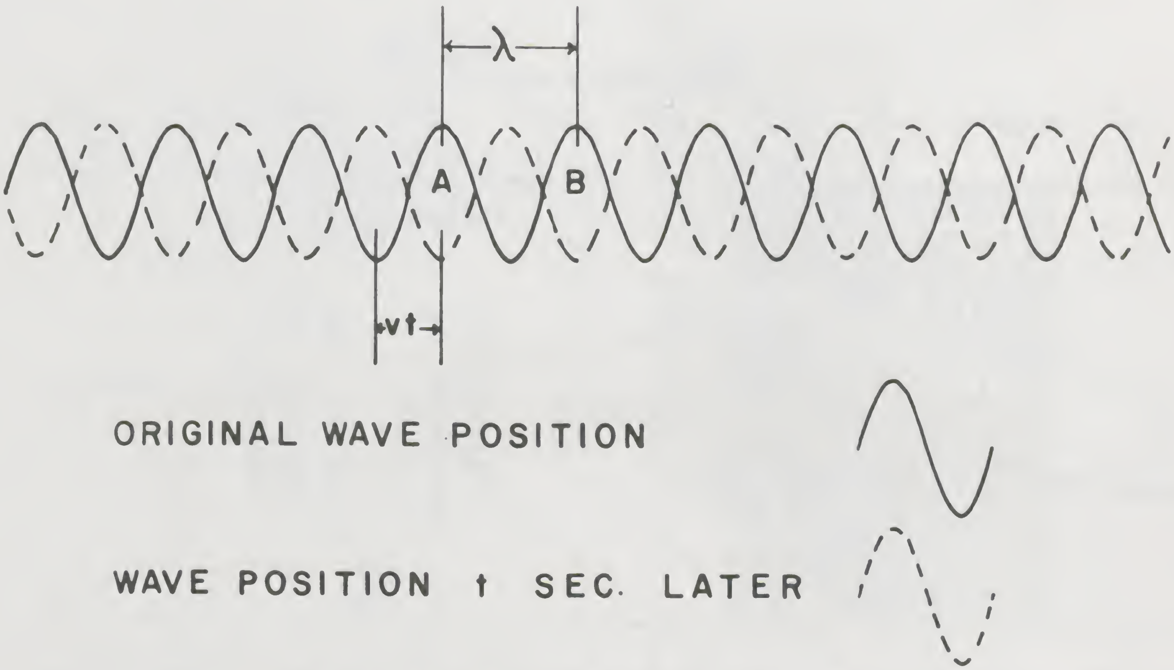


Figure 6. The wave moves to the right with speed $v$.

*The upper limit is different for different people. For a particular individual, this limit becomes lower as the person gets older.

$$f = v/\lambda = (3 \times 10^8 \text{ m/s})/(4 \times 10^{-7} \text{ m})$$
$$= .75 \times 10^{15} \text{ s}^{-1} = 7.5 \times 10^{14} \text{ Hz}$$

This frequency is enormous when compared to the highest frequency of sound which can be heard.

Waves that are visible to us as light are just a small part of a much larger class of waves, called *electromagnetic waves.* You are already familiar with the names of many other types of electromagnetic waves, such as radio waves, microwaves, x-rays, ultraviolet rays, and infrared radiation. All of these electromagnetic waves travel at the same speed in empty space. They are identical except for wavelength and frequency. The complete range of wavelengths or frequencies of electromagnetic waves is called the *electromagnetic spectrum* (Figure 7).

The human eye is sensitive to only a small part of the electromagnetic spectrum, the *visible spectrum.* Each different wavelength within the visible spectrum is a different color. Thus, to compare more accurately the yellows in the different *spectra* (plural of spectrum) you've observed, you need to measure the wavelengths of the light in the two cases.

**Problem 1.** Find the wavelength for electromagnetic waves whose frequencies are

a. 60 Hz (the same frequency as household electric power);

b. 560,000 Hz (560 kHz) (this is a frequency near the low end of the standard broadcast band);

c. $7 \times 10^{14}$ Hz (light).

**Problem 2.** Find the frequency of electromagnetic radiation whose wavelength is

a. 400 nm

b. 600 nm

c. 90 nm

d. 1 m

e. 100 m

For each, indicate the kind of radiation it is.

**Problem 3.** Of the following wavelengths, which are those of visible light?

120 nm 5000 nm 10 nm 1000 nm
500 nm 430 nm 650 nm

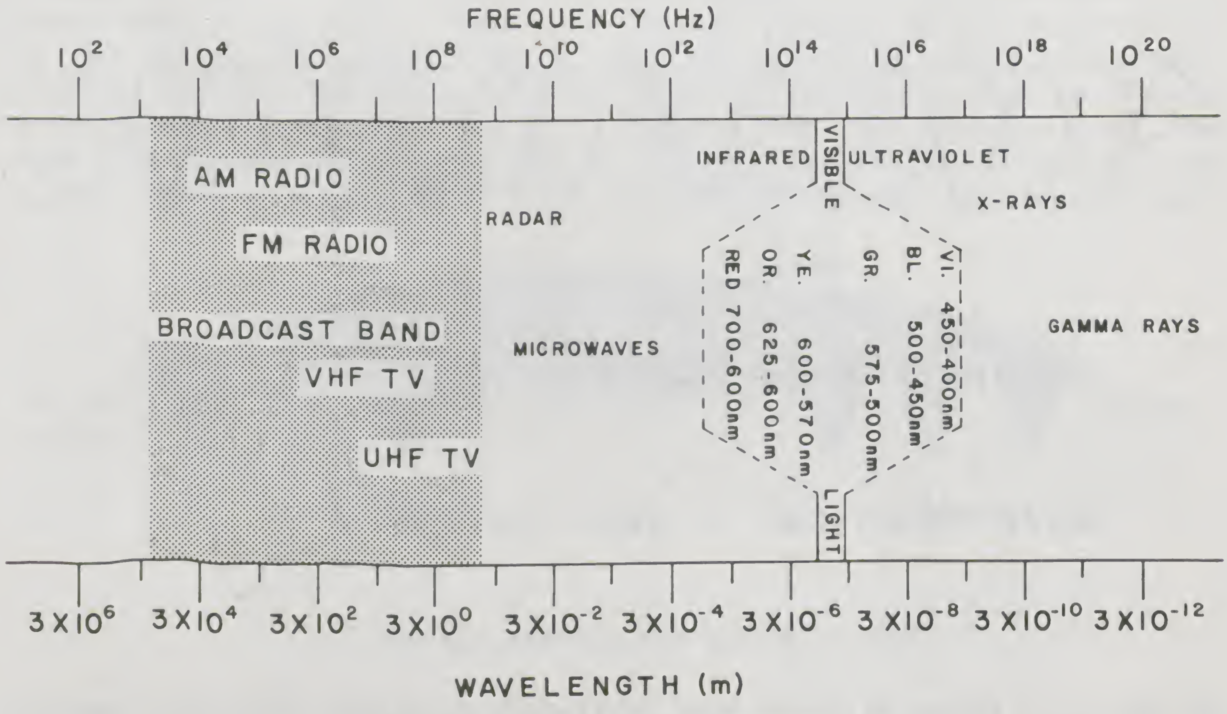


Figure 7. The electromagnetic spectrum.

**Question 4.** Can you explain why there are not separate images of the source in a continuous spectrum?

When you looked through the diffraction grating at the clear end of the fluorescent lamp, you saw several images of the lamp, displaced to each side of the actual lamp. The distance by which each of these *images* appears to be displaced from the actual lamp depends in a simple way on the wavelength of the light. A diffraction grating can therefore be used to measure the wavelength of light in the visible spectrum.

**EXPERIMENT A-3. Measurement of Wavelength**

A. Place the sheet of cardboard with the narrow rectangular slit in front of the bunsen burner and a meter stick in front of the slit, as shown in Figure 8. Light the burner and again place some rock salt on it. Look through your grating and read the distance from the actual flame to the first image of the slit on each side. (You may need some help from a partner who slides his finger along the meter stick until you tell him that it is at the same position as the yellow image.) Do the same for the second image on each side, and fill in Table I. Also measure the distance, $L$, from the slit to the grating and enter this into Table I.

You can now relate your readings on the meter stick to the wavelength of the light. A formula for the wavelength is

$$\lambda = \frac{dx}{n\sqrt{L^2 + x^2}}$$

In this formula $d$ is the distance between the grooves on the diffraction grating (see the optional discussion at the end of this section), $x$ is the measured distance along the meter stick, and $n = 1$ for the first image and $n = 2$ for the second image on each side.

For a diffraction grating with 13,400 grooves per inch, $d$ is approximately 1900 nanometers ($1.9 \times 10^{-4}$ cm). If you have a different grating, get $d$ from your teacher.

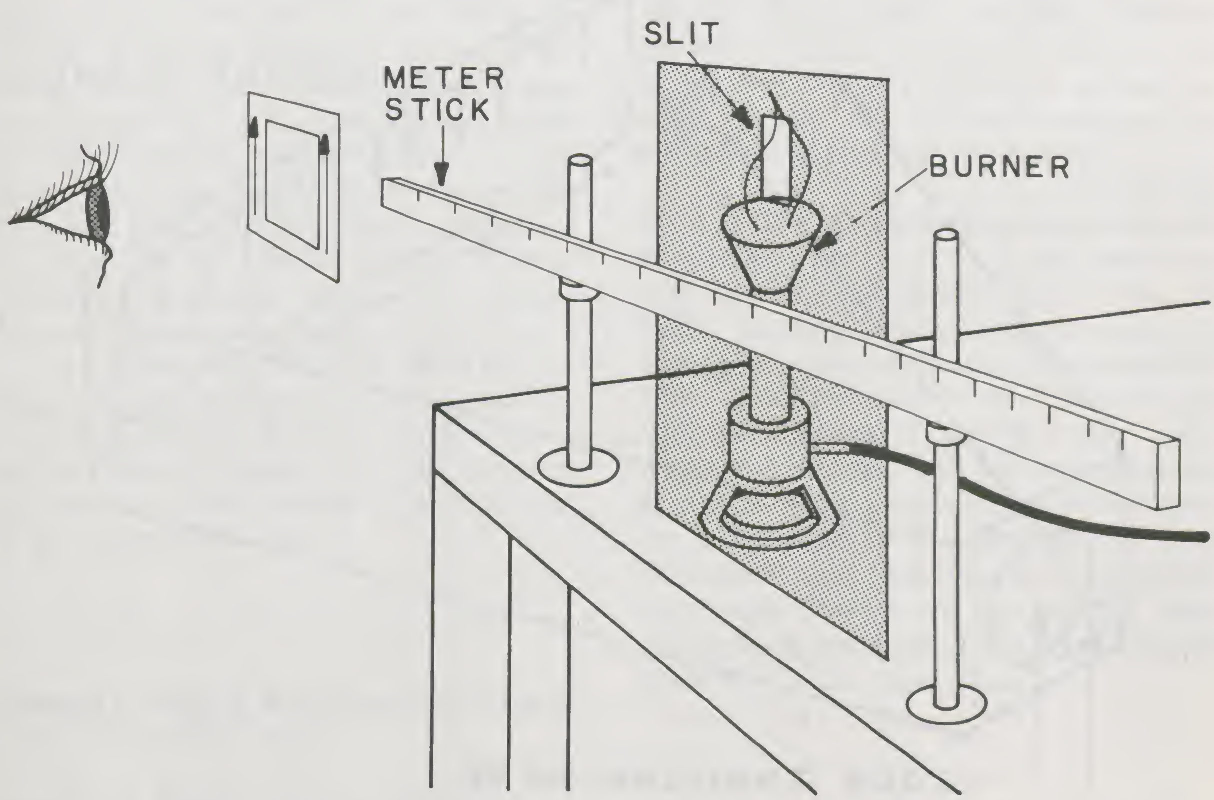


Figure 8. Arrangement for the grating experiment.

Table I

| | Distance to Image from Slit | | | |
|---|---|---|---|---|
| | Left Side | Right Side | Average of Both Sides | $L$ |
| First Image | | | | |
| Second Image | | | | |

(For more information on gratings, see the optional material at the end of this section.) Using your measurement for $L$ and the average value of $x$, both in the same units, compute the wavelength of the yellow light. Does this lie in about the right region of the visible spectrum? Do you get about the same answer for $n = 1$ as for $n = 2$?

B. These ideas are the basis for the construction of an instrument called a *grating spectrometer.* Instead of using a meter stick, the ready-made spectrometer (Figure 9) consists of a piece of diffraction grating, a slit, and a wavelength scale. You can read the wavelength of any spectral line directly from the scale in hundreds of nanometers.

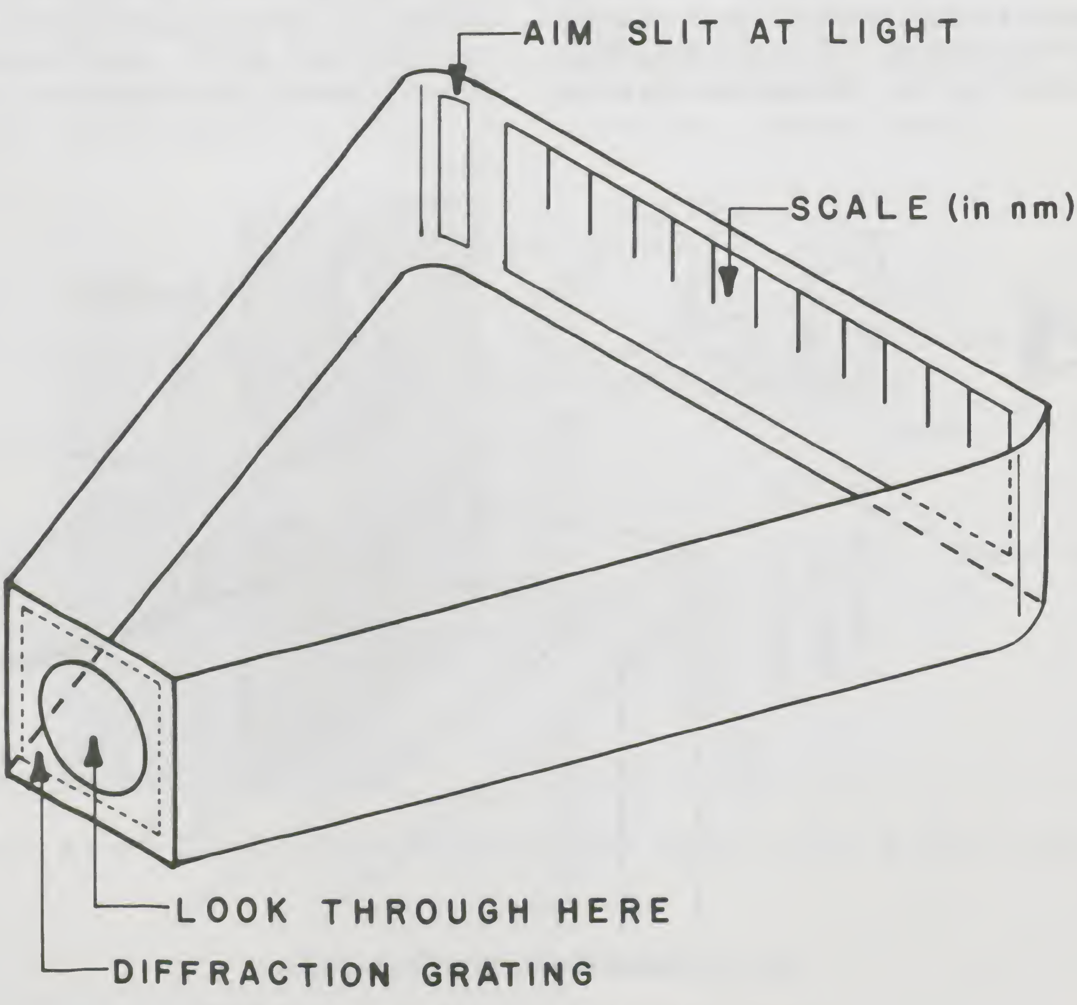


Figure 9. A cutaway view of a typical grating spectrometer.

Use the grating spectrometer to measure, as carefully as you can, the wavelengths of the bright yellow line in the sodium spectrum. Then go back and measure all of the lines you can see in the spectrum of the clear end of the fluorescent lamp. Is the wavelength of the yellow line from the flame the same as that of the yellow line in the fluorescent spectrum? Also use the spectrometer to measure the upper wavelength limit (red) and the lower wavelength limit (blue) of the incandescent lamp spectrum.

---

You have now measured two different spectra, and you have observed a third (from the incandescent lamp). Do other lights have other spectra? You need to observe as many light sources as you can.

### EXPERIMENT A-4. A Take-Home Experiment

Tape your diffraction grating to a 3 X 5 index card so that it may be folded back against the card for protection in your pocket. Take it home with you and observe all the lights you can. You can describe their spectra on the index card for future reference. Look at street lights, car head lights, household lamps, neon signs, strobe lights, black light, stars, etc. Note any similarities or differences you see. (Do not look at the sun. Without pain or other warning signals, direct sunlight can cause permanent damage to your eyes.)

It is likely that most of the lights you are able to observe outside the laboratory will be either fluorescent or incandescent. Their spectra will either be continuous, or will exhibit a line structure similar to that of the clear end of the fluorescent lamp.

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**Question 5.** What is the purpose of the slit in front of the sodium flame?

**Question 6.** Would you expect moonlight to have a continuous spectrum or a line spectrum? Why?

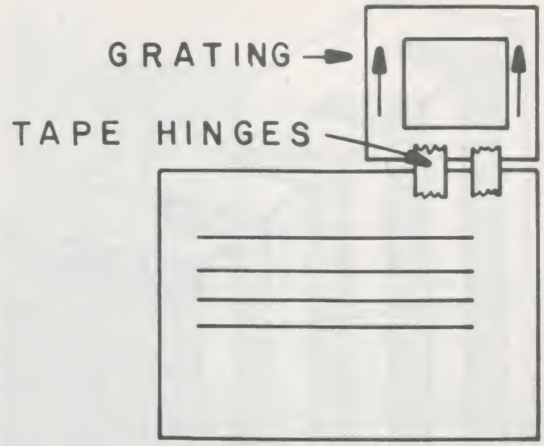


Figure 10. Arrangement of diffraction grating and 3 X 5 index card.

### THE DIFFRACTION GRATING (Optional)

In Experiment A-3, you found that meter-stick readings could be related to the wavelength of the light by means of a formula. Here, and in Experiment A-5, we shall look at where that formula came from.

A diffraction grating consists of a sheet of transparent material* (such as glass or plastic), with a large number of narrow grooves or ridges, separated by narrow flat spaces (Figure 11). The ridges or grooves are evenly spaced and we shall designate the distance of separation by the letter *d*.

As light passes through a narrow flat space, it does not all continue along the same straight line, but much of it gets bent away from the original direction. This bending of light as it passes through a narrow opening is what is meant by *diffraction*. The combined effect of all the narrow "slits" between the ridges or grooves is that the light going through is canceled out except for certain directions. The images you have been observing are formed by the light bent in these "preferred" directions (that is, it is not canceled out). Before we can specify these directions in more detail, we must consider other parts of the arrangement.

*Such a grating is properly called a *transmission diffraction grating*; there are also *reflection* diffraction gratings.

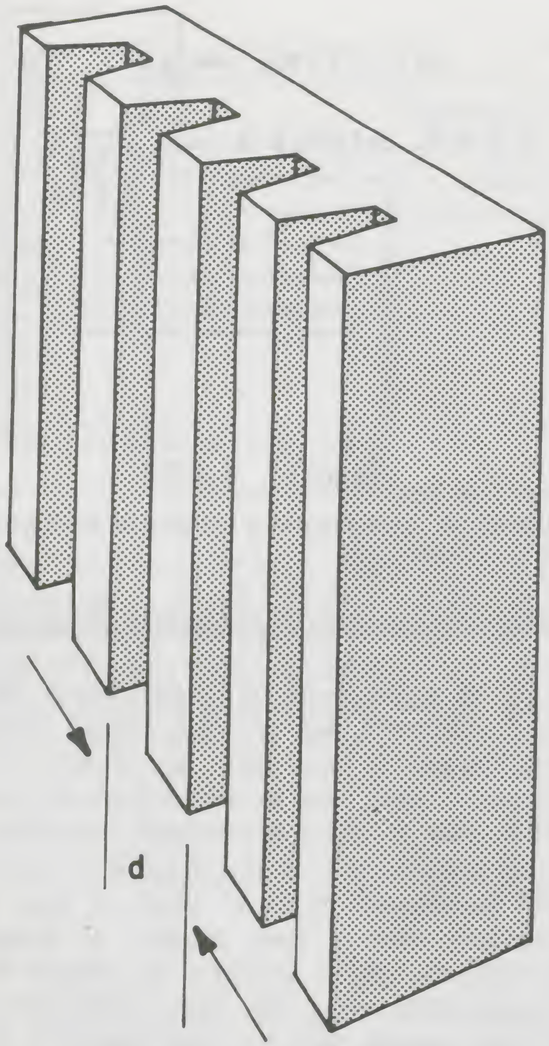


Figure 11. The grooves or ridges of a diffraction grating.

Recall that you saw *repeated* images of the sodium flame (or slit) on each side of the actual flame. Let us number the images on either side, using 1, 2, 3...., starting with the image nearest the flame (Figure 12). Let $n$ represent this number. The $n = 1$ spectrum is called the *first order* spectrum; the $n = 2$ spectrum is *second order;* etc. Each higher-order spectrum is dimmer than the preceding one. How many orders are visible depends on the quality of the grating and the sensitivity of the eye.

The formula giving the direction in which light of a given wavelength is bent to form an image is

$$n\lambda = d \sin \theta \tag{2}$$

If the spectrum consists of more than one line, then Equation (2) gives the direction for each colored image. For each value of $\lambda$ there is a value of $\theta$ for the image of that color.

The spacing $d$ is determined by the construction of the grating and is usually known. In our case, we can use our measurements to calculate $d$.

**EXPERIMENT A-5. The Diffraction Grating (Optional)**

Set up the apparatus again, as shown in Figure 13, with a slit in front of the burner.

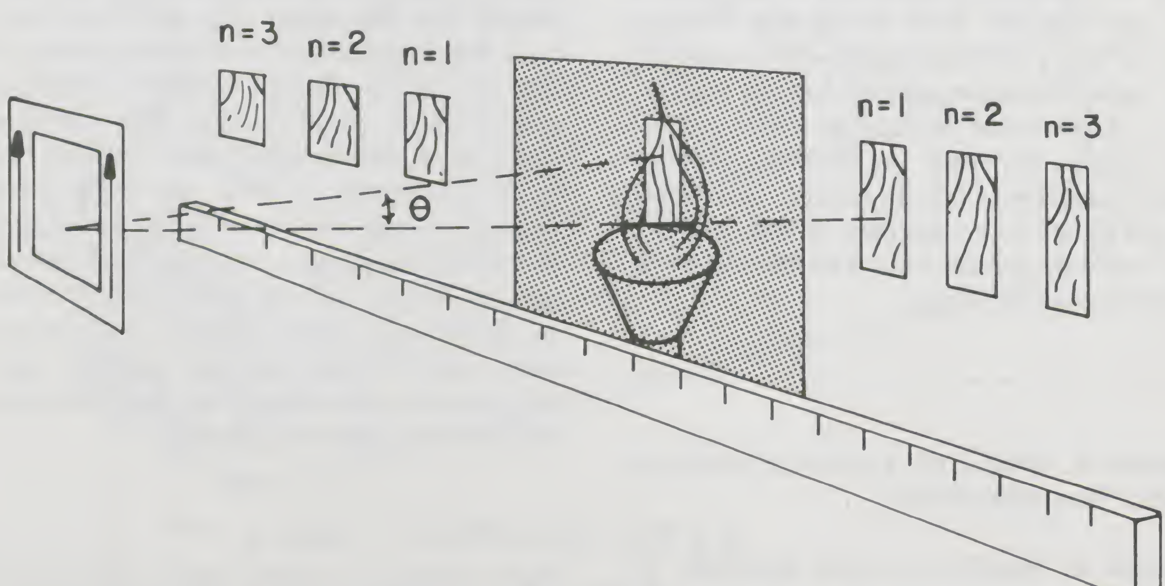


 Figure 12. Numbering of the repeated images. The angle $\theta$ is measured from the slit itself to any one of its images.

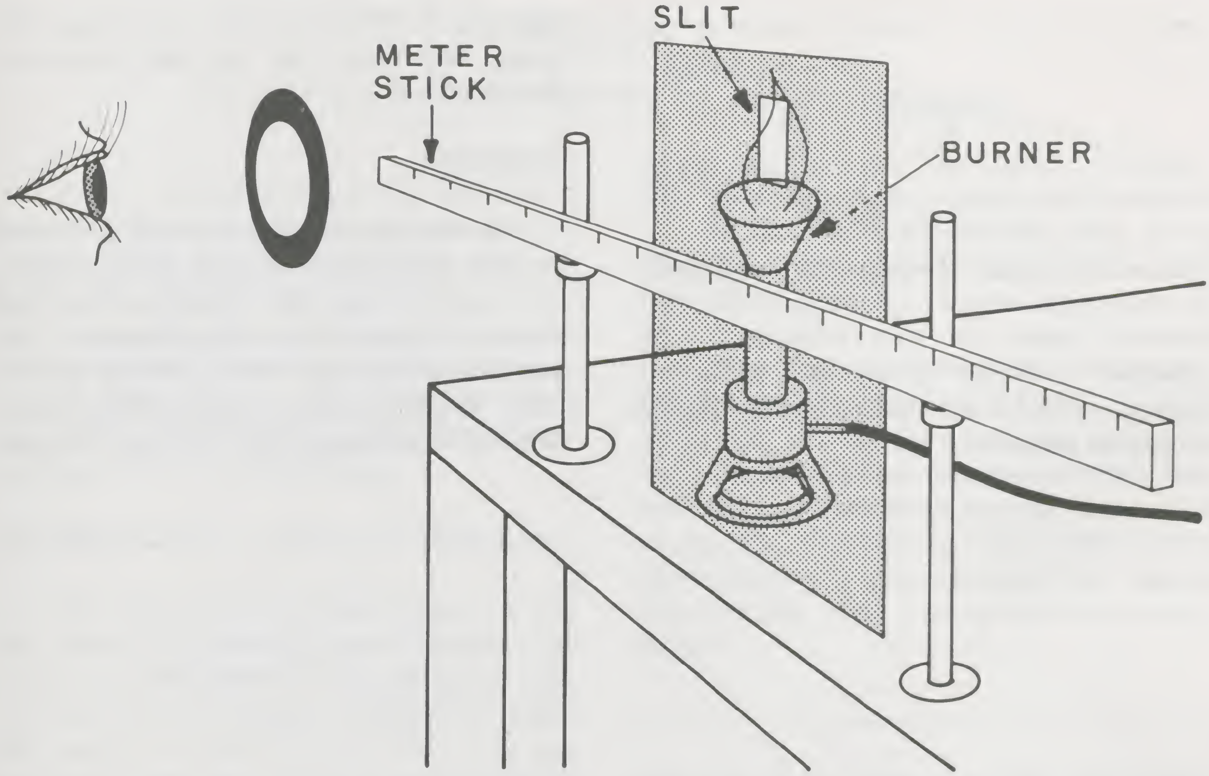


Figure 13. Arrangement for the grating experiment.

Record the distance ($L$) between the grating and the meter stick, just in front of the slit. This time use a diffraction grating with different groove spacing. Then record the distance ($x$) from the slit to each of its yellow images along the meter stick, using Table II.

For a particular image, the angle $\theta$, which is the angle by which the image is displaced from the slit, can be computed from the measurements (see Figure 14).

From trigonometry, we have

$$\tan \theta = x/L \qquad (3)$$

You measured the value of $\lambda$ with the grating spectrometer in Experiment A-3. By substituting the average value of $x$ for the first image ($n = 1$) into equation (3), find $\theta$. Then, using Equation (2), find $d$. Repeat this procedure for each value of $n$. Are your values of $d$ the same for each case? Find the average of $d$ from these values.

Knowing the value of $d$, you can compute unknown wavelengths (from other sources) if $x$ is measured. See if you can figure out a way to measure the wavelength of the green line of the fluorescent lamp spectrum.

Table II

| $n$ | $x$ (left side) | $x$ (right side) | $x$ (average value) |
|---|---|---|---|
| 1 | | | |
| 2 | | | |
| 3 | | | |

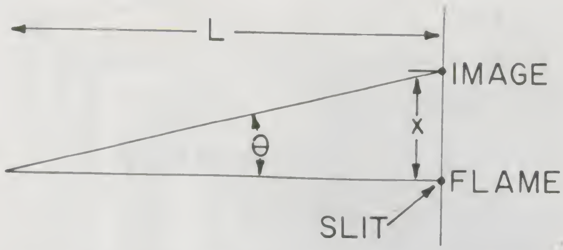

Figure 14. Top view of experimental arrangement.

**Question 7.** Which of the two diffraction gratings you have used has the greater spacing *d* between grooves?

**Question 8.** Which of the two spreads the images out more, the one with greater or smaller *d*?

## SUMMARY

In this section of the module you studied light from various sources, using a *diffraction grating.* You were introduced to such terms as *continuous spectrum, frequency,* and *wavelength.* You then used a *grating spectrometer* to measure wavelengths and learned to correlate wavelength with the color of light.

# SECTION B

## The Relationship of Spectra to Atomic Structure

In the previous section you observed a few spectra and became acquainted with the grating spectrometer. In this section you will make more detailed measurements of a variety of spectra and you will also see how to relate these spectra to the changes that take place within atoms as they emit light. This discussion will, in turn, help in the understanding of the fluorescent lamp.

**EXPERIMENT B-1. Spectra of Other Gases**

A. Clip one of the discharge tubes into the clips provided inside the power supply box (Figure 15). (CAUTION: These lamps are fragile. Also, do not touch the apparatus when the lamp is turned on or you will get shocked.) Turn the lamp on and carefully observe the light. What is its color? Then look at the lamp through the diffraction grating and record what you observe, using crayons.

Repeat this procedure for each of the discharge tubes available. If two lamps emit light which appears to be the same color, do they necessarily have the same spectrum?

From the observations you have made, would you say that any of the discharge lamps give off just the same color of light as the clear part of the fluorescent tube? Measure the spectral lines of the one which seems to be the best match, using your grating spectrometer. Is the spectrum the same as that for the clear part of the fluorescent lamp?

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**Question 9.** Is there any clear relationship between the apparent color of a lamp and its line spectrum?

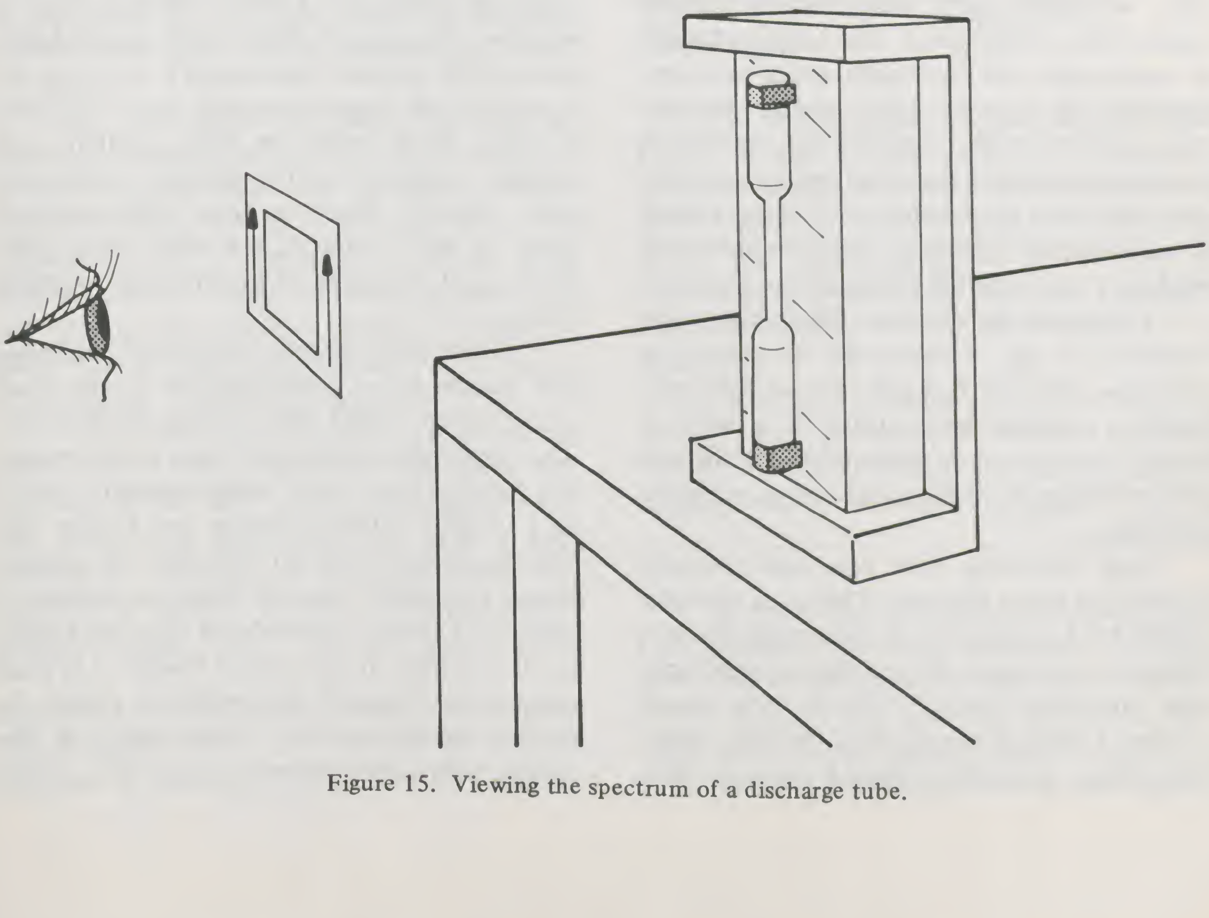

Figure 15. Viewing the spectrum of a discharge tube.

**Question 10.** Why is a slit not needed when viewing the spectrum of a discharge tube?

The fact that several different gases are seen to emit light with quite different spectra leads us to inquire more closely into the nature of the gases themselves. What is a gas made of? To answer this question, we need to look more closely at the structure of matter.

## THE STRUCTURE OF MATTER

Imagine that you could enlarge a substance, say a piece of gold, 100 million times. You would see the gold to be composed of a regular arrangement of small, fuzzy looking spheres, each about 1 cm in diameter. These little spheres are the smallest pieces of gold which exist. They are called *atoms* of gold. Every item (solid, liquid, or gas) which is identifiable as gold is built of gold atoms.

There are about 100 different kinds of atoms, which are the building blocks of *all* materials. A substance which is built of one single kind of atom is called an *element*, so there are about 100 different elements. All other substances are combinations of elements, called *compounds*. The building blocks of compounds are *molecules* which are combinations of two or more atoms. Because atoms are too small to see directly, everything that is known about them has been learned by observing how they behave in a large variety of experiments. Many of those experiments involved a study of the light emitted by gases.

To answer the question which began this discussion, a gas is composed of atoms (or molecules, if it is the gas of a compound). Whether a particular material is a solid, a liquid, or a gas at a given temperature and pressure depends on the nature of its atoms or molecules.

Each discharge tube you used contains the gas of a single element. That is, it contains a single kind of atom. Each tube gives rise to a different spectrum. If you had a discharge tube for every material known, you would observe a different spectrum for each tube. The unique spectrum emitted by each tube can be used as a "fingerprint" to identify the particular gas in that tube. No two different materials have ever been found to emit the same line spectrum. Thus, if you have found a discharge tube whose spectrum matches that of the clear part of the fluorescent, both must contain the same element. What element is in the fluorescent lamp?

**Question 11.** Can you identify, from your own observations, other types of lamps which have the same line spectrum as that of the fluorescent lamp?

It seems that a particular line spectrum belongs in a unique way to the atoms which produce it. The beautiful arrangement of lines seen through the diffraction grating yields much insight into the structure of the atom. Before we can gain this insight, we need to be aware, in a general way, of what an atom is.

## ATOMIC STRUCTURE

What are atoms made of? There are many clues that, in all materials, there are two different kinds of electric charge, labeled positive and negative. Since most materials are electrically *neutral* (uncharged), and yet all materials will respond in some way to electric charges, there must be both positive and negative charges in equal amounts within the same material. Since an atom is the smallest piece of any material, the atom must itself have equal amounts of positive and negative charges.

Other experiments show that the positive charge in an atom exists in a very small central core, called the *nucleus* of the atom. The negatively charged electrons move around the nucleus in a much larger region of space, in a roughly spherical region (see Figure 16). The electrons move so fast that the mental image physicists have of them is a kind of blur, or a fuzzy, smeared-out *electron cloud*, as it is called. If one could enlarge a typical atom to a diameter of about one meter, its nucleus would still be a tiny speck in the center, with a diameter of between $10^{-5}$ meters

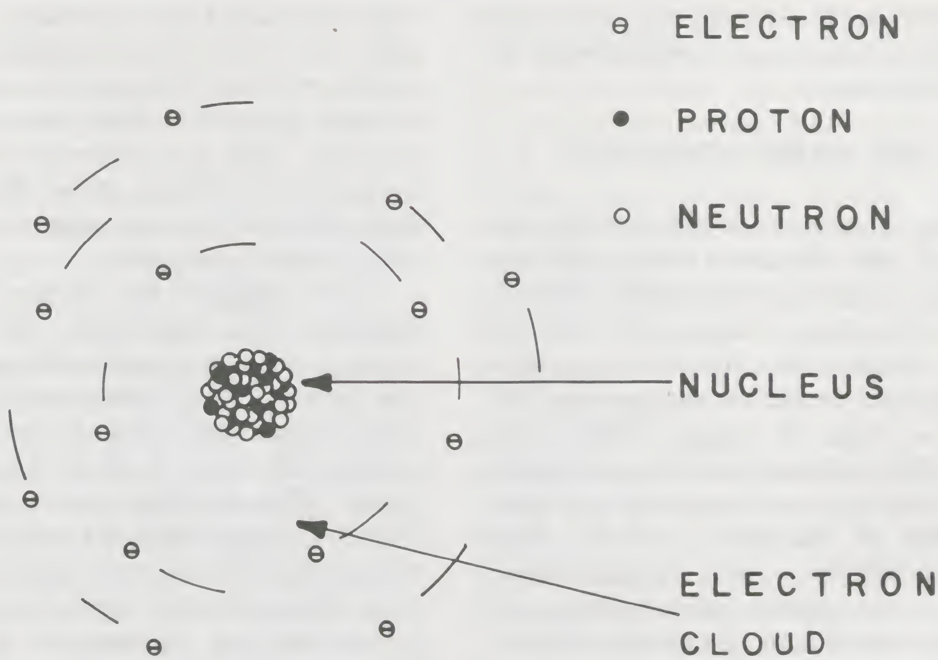


Figure 16. Our mental picture of an atom.

and $10^{-4}$ meters. (Or perhaps, if the atoms were enlarged to the size of the Houston Astrodome, the nucleus would be about the size of a golf ball at its center.) The atom is held together by the electrical force of attraction between the positively charged nucleus and each of the negatively charged electrons.

The nucleus of the atom consists of positively charged particles called *protons* and neutral particles called *neutrons.* Protons have a positive charge which is the same size as the negative charge of the electron. So an atom which contains equal numbers of electrons and protons has zero net charge; it is *electrically neutral.* The uncharged neutrons have no effect on the charge of the atom.

The number of protons in the nucleus is called the *atomic number* of the atom (or element). When the atom is electrically neutral, the atomic number is the same as the number of electrons in the atom. Each element has a different atomic number. Atomic numbers range from one to about 100.

**Question 12.** The carbon *nucleus* has a diameter about $6.9 \times 10^{-13}$ cm, whereas the carbon *atom* has a diameter about $1.5 \times 10^{-8}$ cm. If we want to build a model of a carbon atom with its nucleus the size of an orange (diameter 8.0 cm), what would be the diameter of the space for the electron cloud?

**EXPERIMENT B-2. The Spectrum of Hydrogen**

The simplest of all atoms is the hydrogen atom, whose atomic number is one. Its nucleus usually has one proton and nothing else. A single electron makes up its electron cloud. One might expect that the simplest of all elements would also have an especially simple spectrum. Let us take a look at its spectrum.

Using the hydrogen discharge tube and your grating spectrometer, observe again the visible lines of the hydrogen spectrum. You should be able to observe four lines.

---

As you can see, there is nothing especially simple about the spectrum. The fact that a combination of a single proton and a single electron can give rise to such a complicated spectrum suggests that the spectrum is not related in a simple way to the number

of electrons in the cloud. To understand better what is happening we must discuss the energy of the atom.

## ENERGY AND ATOMIC STRUCTURE

Energy is an essential ingredient for the creation of light. Electrical energy and heat energy were used to generate light from the lamps and the flame, respectively. This observation, added to the idea of conservation of energy, leads to the conclusion that light must be a form of energy. This is not surprising. We all know that heat accompanies light from the sun and from an incandescent lamp. Most of this heat is in the form of *infrared radiation,* electromagnetic waves with frequencies slightly lower than those of visible light (see Figure 17). Radio stations, transmitting radio waves, express their output in watts, which measure the rate at which energy is emitted. Microwaves can be used to cook food in a new type of oven. All forms of electromagnetic waves carry energy.

Two common forms of energy are kinetic and potential. *Kinetic energy* is the energy that a body has because of its motion. The *potential energy* of a body depends on the position or the shape of a body. Also if work is done on a body, its energy increases. In turn, a body can *do* work by giving up some of its energy. Work is a measure of the amount of *energy change* in these cases.

For an atom, we can subdivide the energy in a still different way. The atom as a whole can be moving and thus it can have kinetic energy. Its location (for example, near another atom) might give it potential energy. We can also speak of the energy that is *internal* to the atom. Since the atom consists of a nucleus and a cloud of electrons, we might expect that each of these major parts can have energy of its own. This is found to be the case. The term *nuclear energy* refers to the energy of the nucleus. It depends on the structure of the nucleus–the number of protons and neutrons and their relationships.

The term *atomic energy* properly refers to the energy of the electron cloud. Unfortunately, *atomic energy* is often used to refer to what was called *nuclear energy* above. To avoid confusion, we shall use the term *energy of the atom* to refer to the energy of the electron cloud. In further discussions of the atom, we shall be concerned only with the energy of the electron cloud. We will therefore often use the single word *energy* instead of the longer expression.

The energy of the electron cloud for a particular atom depends on the number of electrons it has and on the size and shape of the electron cloud. Unless something has put energy into them, most of the atoms of a sample will stay at their lowest possible energy. This is called *ground-state energy.* Work (or energy input) is required to increase the energy of the atom above the ground-state value. Generally, this makes the cloud larger, by moving one or more of the electrons farther from the nucleus. When the energy has thus been increased, we say that the *atom is in an excited state*, or the *atom is excited.*

A case of special interest is the one in which enough energy is given to the atom to remove one or more electrons completely from the electron cloud. We then say that the atom has been *ionized.* The remaining atom is no longer electrically neutral. It has a net charge equivalent to that of a proton. Such a charged particle is called an *ion.* The removed electron and the remaining charged atom together are called an *ion pair.*

**Question 13.** The atomic number of mercury is 80. Is the atomic number of an *ionized* mercury atom still 80?

**Question 14.** Does an ion pair have a higher energy or a lower energy than the normal atom before ionization?

## ENERGY LEVELS FOR AN ATOM

Are there any rules governing the amount by which the energy of an atom can be increased? Experiments have shown that the energy cannot change by just any arbitrary amount. For a particular atom the possible amounts by which the energy can change are fixed.

You are probably accustomed to thinking that the energies of ordinary bodies (a stone lying on a hillside, a baseball moving through the air, an automobile rolling along a highway, etc.) can have *any* value. Kinetic energy is related to the speed of motion and the speed can have any value we choose. Potential energy is related to position (as in the case of a stone on a hillside) and it depends on the height above some starting point. This height can be anything we choose.

In the case of the energy of the electron cloud surrounding an atom, the value of the energy is *not* anything we choose; it happens that only certain values are possible. It is as if the hillside consists of a series of ledges with no gradual sloping between, and the stone must lie on one of the ledges. In Figure 17, such a hillside is shown and, drawn to scale, an *energy-level diagram* corresponding to it.

We use the term *energy level* to designate a possible value of the energy of the electron cloud. The diagram of possible energies is then called an *energy-level diagram.* In Figure 18, an energy-level diagram for some imaginary stone is shown. The lowest energy, called the *ground-state energy,* is labeled $E_0$ and the other values are $E_1$, $E_2$, $E_3$, etc. In the diagram the separation of any two lines represents the amount of energy difference between the two levels (just like the "ledge diagram" for the hillside). Any one stone can only have one energy at a given time, but the diagram shows the possibilities.

An energy-level diagram not only gives the possible energies, but also the possible energy *changes.* Thus, for Figure 18, possible energy changes are $(E_1-E_0)$, $(E_4-E_3)$, $(E_2-E_5)$, etc., in any condition. In all cases the possible energy changes are differences between values for various levels.

When the energy of an atom changes, we say that the atom *undergoes a transition.* The atom which has a certain energy is said to be in a certain *state,* so that energy changes are produced by *transitions between energy states.*

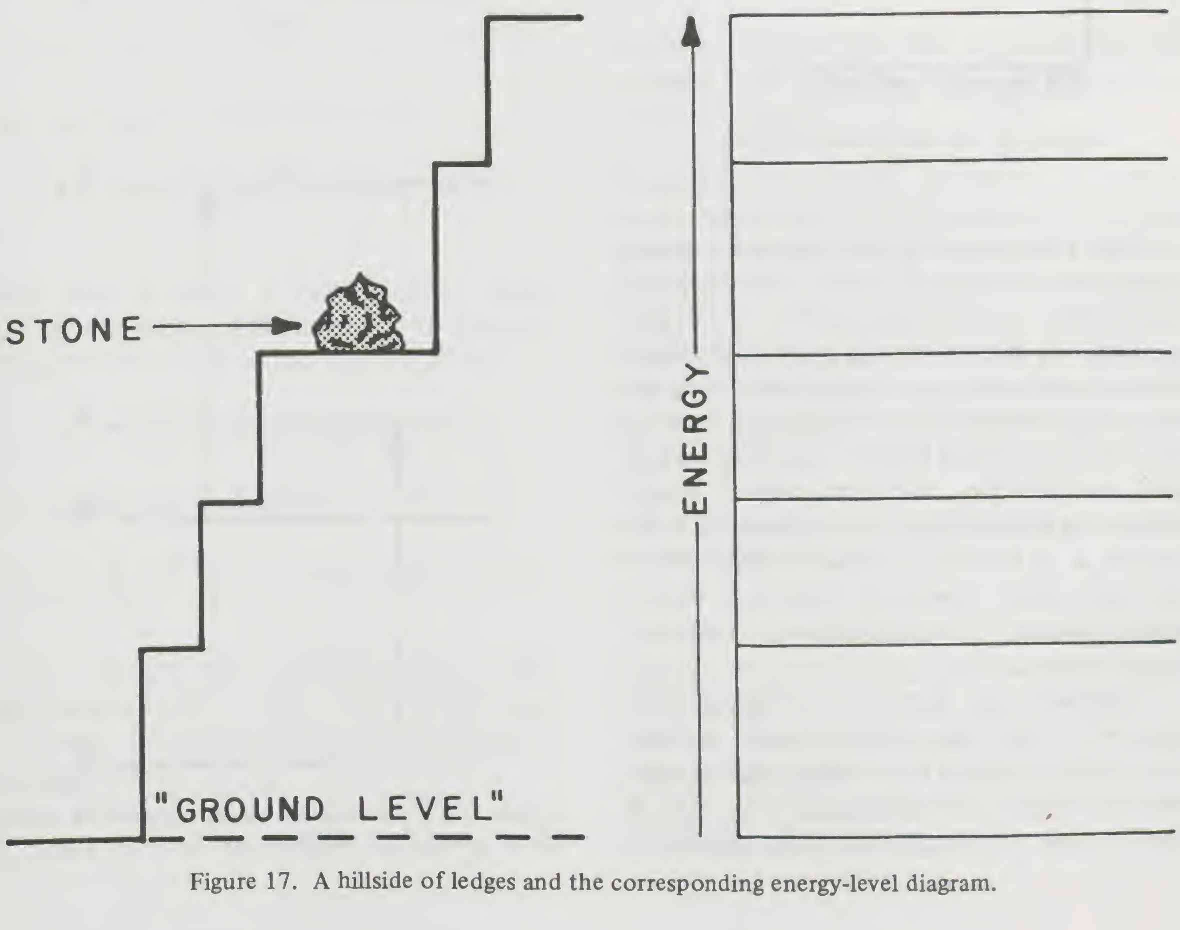


Figure 17. A hillside of ledges and the corresponding energy-level diagram.

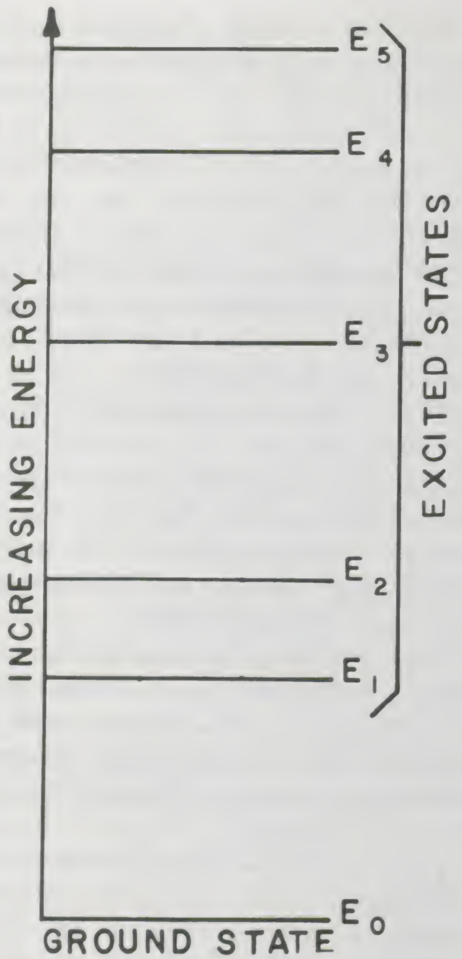

Figure 18. An energy-level diagram.

Let us suppose that an atom can receive an amount of energy $\mathcal{E}$. If the atom is in the ground state at the beginning, the new energy level will be higher by the amount $\mathcal{E}$. If the atom was already in an excited state, then the new energy level will be still higher, again by the amount $\mathcal{E}$. (See Figure 19.) These transitions are possible, of course, only if energy differences (in our case, $E_2 - E_0$ and $E_3 - E_1$) are just $\mathcal{E}$. If the atom has no energy level of the right value, then the transition cannot take place and the energy that is "offered" cannot be accepted.

Similarly, an atom in an excited state can lose energy via a transition to a lower state. If the atom is in a highly excited state (several levels above ground state), then it might return to its ground state directly in one jump, or it might "stop along the way" at various other levels. In any case, the atom usually returns to the ground state very quickly and remains there until it is again *excited* by receiving energy from outside. Let us stress that the energy it gets must be just exactly the right amount to make it jump up one or more energy levels; it won't accept some energy in between.

**Question 15.** Draw an energy-level diagram like that of Figure 18. Indicate at least three possible transitions involving the first excited state, $E_1$.

**Question 16.** If an atom is in the ground state, can it lose energy?

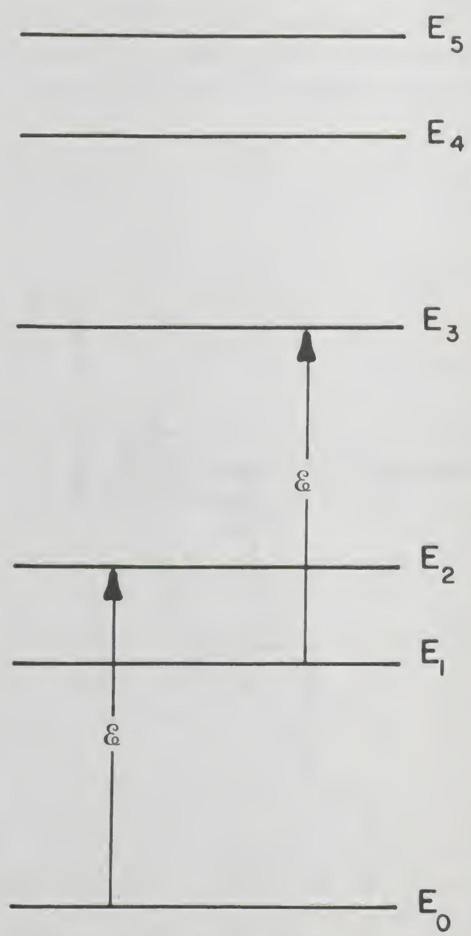

Figure 19. Two possible energy transitions when a given amount of energy is added to the atom.

## ENERGY TRANSITIONS AND LIGHT

The above discussion of the energy of the atom is important to this module because the way an atom gives off light is by *losing* energy in a transition. It *might* seem logical that the amount of energy lost by the atom would determine the brightness (or intensity) of the light emitted. But this is not the case. The amount of energy lost by the atom determines the *frequency* of the light. The relationship is

$$\mathcal{E} = hf \tag{4}$$

where $\mathcal{E}$ is the energy of the transition, $f$ is the frequency of the light, and $h$ is a constant, called "Planck's constant." Energy is measured in joules, frequency in Hertz, and $h$ has the value $6.63 \times 10^{-34}$ joule · second (J · s).

Let us calculate the energy of the transition giving the yellow line of the sodium spectrum, $\lambda = 589$ nm. The frequency is

$$f = v/\lambda = \frac{3.00 \times 10^{8}\ \text{m/s}}{589 \times 10^{-9}\ \text{m}} = 5.09 \times 10^{14}\ \text{Hz}$$

Then the energy of the transition is

$$\begin{aligned} \mathcal{E} &= hf = (6.67 \times 10^{-34})(5.09 \times 10^{14})\ \text{J} \\ &= 3.37 \times 10^{-19}\ \text{J} \end{aligned}$$

Since this is such a tiny number when expressed in joules, it is customary to measure energies of the atom in *electron volts* (eV):

$$1\ \text{eV} = 1.60 \times 10^{-19}\ \text{J}$$

Then for the above example,

$$\begin{aligned} \mathcal{E} &= 3.37 \times 10^{-19}\ \text{J} \times (1\ \text{eV})/(1.60 \times 10^{-19}\ \text{J}) \\ &= 2.11\ \text{eV} \end{aligned}$$

Energy changes that produce visible light range from about 1.8 eV to 3.1 eV. In order to simplify the calculations, we can combine the operations as follows:

$$\mathcal{E} = hf = hv/\lambda$$

If $\lambda$ is measured in nanometers and $\mathcal{E}$ in electron volts, then $hv$ has the value 1240 eV· nm. Thus

$$\mathcal{E}\ (\text{in eV}) = \frac{1240\ \text{eV·nm}}{\lambda\ (\text{in nm})} \tag{5}$$

The energy-level diagram for atoms of any one element is different from that for atoms of any other element. The transitions are then also different. This fact, in turn, accounts for the different observed spectra of different elements.

Earlier a question was raised about the intensity or brightness of light emitted by atoms. The brightness of a particular spectral line depends, not on the energy of the transition involved, but on the *number of atoms per second* undergoing that particular transition. All the spectra we have studied involve billions of atoms undergoing the same transitions. The radiation from a single atom is so weak that it can be observed only under very special conditions.

**Problem 4.** Compute the frequency of light emitted by a transition of energy $\mathcal{E} = 5$ eV. Is this light visible?

**Question 17.** Discuss, in terms of energy levels, what must be taking place as hydrogen (a one-electron atom) gives rise to a complex spectrum.

**Question 18.** An electron and a proton come together to form a hydrogen atom in the ground state. Why would you expect radiation to be emitted in the process?

## WHAT SPECTRA TELL US ABOUT ENERGY LEVELS

We now have a "tool" for determining energy levels of atoms: the line spectrum of the atom. Knowing that each frequency of emitted light corresponds to an energy change, scientists can "piece together" the energy-level diagram of an atom from the entire line spectrum of atoms of that element.

**EXPERIMENT B-3. Analysis of Mercury Spectrum**

A. From your data for the wavelengths of the mercury spectrum, compute the energy change of the transition that must have taken place to give that wavelength of light. Express these energies in electron volts, using Equation (5).

The energy-level diagram of mercury is too complex to be pieced together from your measurements. Instead, we will give you a simplified energy-level diagram for mercury and you may verify that the observed lines do represent possible transitions.

The simplified energy-level diagram for mercury appears in Figure 20. The lines in the visible spectrum are caused by transitions between a variety of energy levels. Check your measurements of the mercury spectrum against those transitions labeled *A, B, C,* and *D* on the diagram. What are the colors represented by each of these transitions?

**Problem 5.** What is the energy of the transition that gives rise to the blue-green line of the hydrogen spectrum, wavelength 486 nm?

As you can see from the energy-level diagram, many transitions are possible that cannot be seen by using the simple grating spectrometer. More information can be obtained by using a different method.

B. In this experiment you are to use a camera and the diffraction grating which is mounted in glass to obtain a photograph of the mercury spectrum. Since details of the procedure will depend somewhat on the camera and the film used, you will need to consult your instructor or make some trial runs yourself. Only some general hints will be given here.

The glass grating can be taped to the front of the camera, covering the camera lens with the grating material. If the lens is a large one, mask it off so that light enters it only

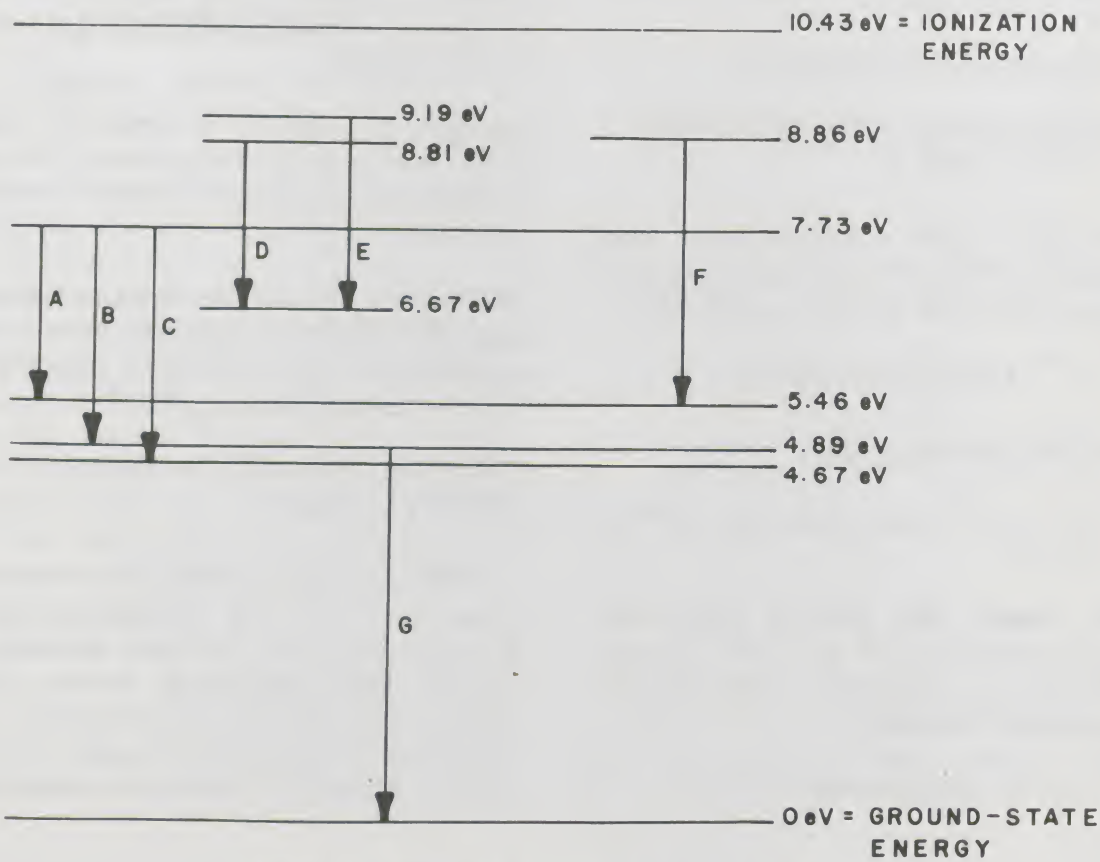


Figure 20. Simplified mercury energy-level diagram.

through the diffraction grating. It would be advisable first to view the spectrum by eye, holding the grating by hand (as you did with the plastic grating earlier). Use a bright object to "mark" the position of the yellow line of the spectrum. A white paper arrow placed under the line works well. (This is necessary because the photograph will not indicate the color of the line.) It is important in this viewing that the grating be in approximately the same position that it will have later when taped to the camera. (If you happen to be using a single-lens reflex camera, you can observe the spectrum through the lens.)

The camera should be aimed more or less toward the yellow line of the spectrum, so that the direct bright light from the discharge tube does not overexpose the film. The exposure time, development process, etc. will depend on the camera and the film. (A Polaroid camera works very well.) Similarly, the distance from the camera to the discharge tube cannot be specified beforehand. For good spread of the spectrum on the film, this camera-to-source distance will probably be between two and five feet.

After the photograph has been developed, identify all the lines which you observed and measured earlier. The photograph should show a few lines that were not visible by eye.

You can measure the wavelengths of the extra lines of the spectrum by analyzing your photograph. The most straightforward way to do this is to make a plot of the wavelengths you know versus the distance between the lines of the photograph. Use a ruler with a millimeter scale to measure the distances, as in Figure 21.

Beginning with the yellow line at zero on the scale, measure the distance (in mm) from the yellow line to each successive line. Then lay out a wavelength scale, running from 300 nm to 700 nm, on the horizontal axis of a piece of graph paper. On the vertical axis, mark off a distance scale in millimeters. Then plot distance versus wavelength for the lines whose wavelength you have already measured with the spectrometer. Draw a straight line that goes approximately through these points (Figure 22).

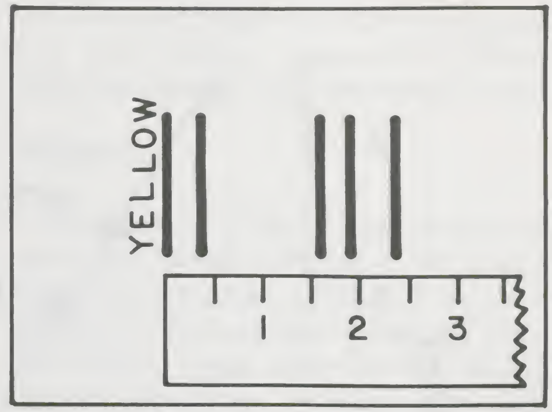

Figure 21. Measuring a spectrum photograph.

You have just calibrated the scale on your photograph. Now, since you know the distances for the unknown lines, you can read their wavelengths from the graph.

Using your *calibration curve*, find the wavelengths of the extra lines in the mercury spectrum. Compute the energies of the transitions that must have produced them. Find these transitions on the energy-level diagram (Figure 20).

In view of the fact that the shortest wavelength most people can see is about 400 nm, does the photograph show any lines invisible to the eye? Compute the wavelength for the transition labeled $F$ in the energy-level diagram (Figure 20). Is this shown on the photograph? Is it visible?

---

The region of the electromagnetic spectrum with wavelengths slightly shorter than those of visible violet light is called the *ultraviolet* region, and the radiation is called *ultraviolet light*. There are many ultraviolet lines in the spectrum of mercury, but most of them are absorbed by the glass of the tube (and of the camera lens). They are therefore not visible on the photograph. If you wanted to photograph them, you would need a special lamp made of quartz, which does not absorb these ultraviolet lines. (You would also need special camera lenses and film which is sensitive to ultraviolet light.)

The most intense line in the entire mercury spectrum comes from the transition

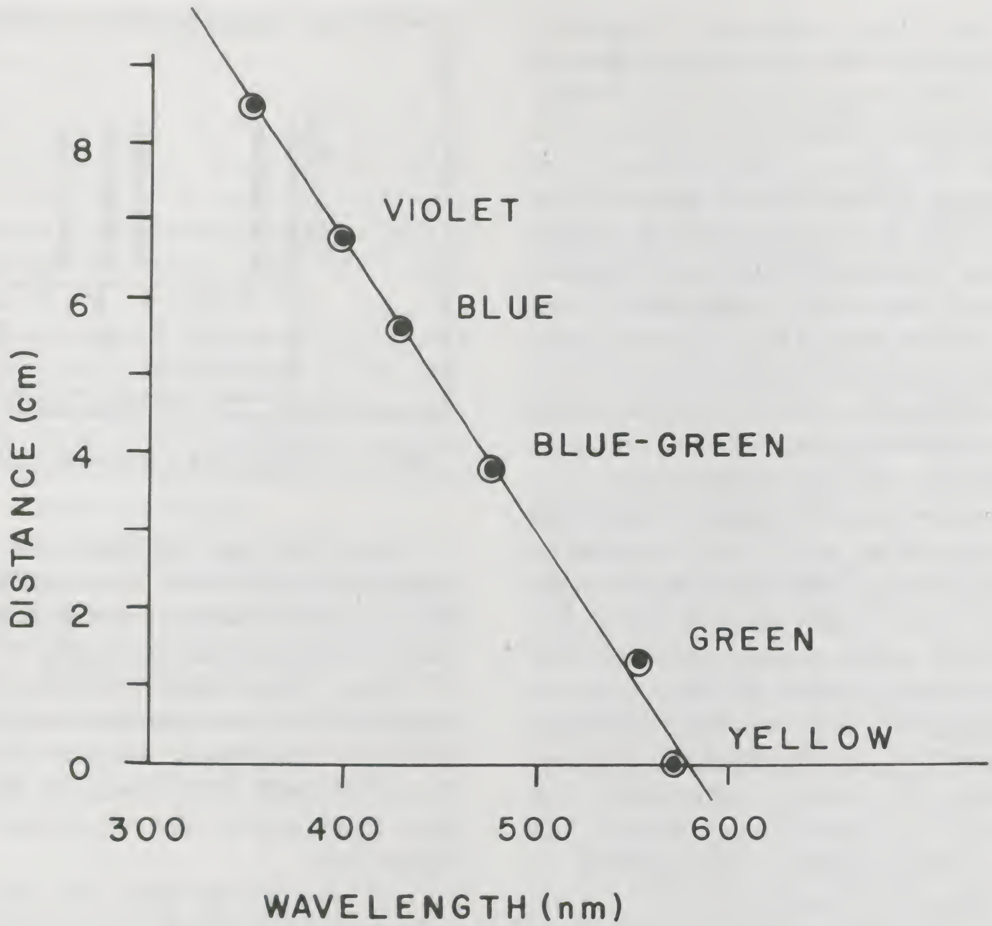

Figure 22. Calibration curve for a spectrum photograph.

labeled $G$ in the energy-level diagram. (Transitions down to the ground state are common for *all* excited atoms.) The wavelength of this line is 253.7 nm. Although it will not pass through the tube, it is *this* line that is most important in the operation of the fluorescent lamp. The connection will become clearer in Section C of this module.

You may have noticed, in Figure 20, that no transition is indicated from the lowest excited state ($E_1$ = 4.67 eV) to the ground state $E_0$. This is not an oversight but rather, it reflects the important point that not all of the transitions which you can draw on an energy-level diagram are actually allowed by nature. The transitions which *do* occur depend not only on the energy levels of the atom but also on other properties of the electron cloud. These properties will not be discussed in this module.

## A LOOK AHEAD

It was stated earlier that atoms emit light (or other radiation) as they *lose* energy by undergoing transitions to *lower* energy states. Ordinarily atoms are found in the *ground state.* In that case they cannot lose energy. Thus, before radiation can take place, something must happen to raise the energy to an excited state—to *excite* the atom. The processes by which this happens will be discussed in the next section.

The general features of the spectra of all atoms are the same. Some are more complicated than that of mercury and they are therefore much more difficult to unravel. But, for all atoms, energy is *absorbed* in definite amounts, raising the energies of the atoms to new levels. When an atom undergoes a transition back to a lower energy level, a particular

frequency of light is emitted. From very careful and often tedious analysis of atomic spectra, we are able to construct energy-level diagrams for other atoms and even for molecules.

It should be emphasized that it is the *electron cloud* that is involved in the transitions giving rise to atomic spectra. For many atoms (including mercury), the visible and ultraviolet spectrum is due mainly to a single electron–the one farthest from the nucleus. For hydrogen, with only one electron (and to some extent for all atoms), energy levels are associated with distances of the electron from the nucleus. Roughly speaking, increasing the energy means taking the electron farther away from the nucleus. With sufficient energy added, the electron gets removed completely from the nucleus (i.e., *ionization* occurs).

Your knowledge of the mercury spectrum gained in this section will (in Section C) be the basis for understanding what happens in the coated half of the fluorescent lamp. We will then discuss commercial fluorescent lamps, blacklights, and other related topics.

## SUMMARY

This section of the module was concerned with more detailed studies of the spectra of gases and the understanding of these spectra in terms of the structure and energy of atoms. You did a visual study of the spectra of various gases in *gas discharge tubes* and you photographed the spectrum of mercury. This photograph showed lines in the ultraviolet region which are not visible with the hand-held diffraction grating or the grating spectrometer. These observations provided the basis for considering *energy levels* of the atom, and *energy-level diagrams* were discussed. Transitions from one level to another were correlated with the line spectrum of mercury.

# SECTION C

## Energy Absorption and Fluorescence

### WHAT HAPPENS TO THE COATING OF THE FLUORESCENT LAMP?

By now you have probably guessed that it is the mercury vapor that is somehow responsible for all the light which comes from a fluorescent lamp. Let's look at this in more detail.

### EXPERIMENT C-1. A Closer Look at the Fluorescent Lamp Spectrum

Carefully repeat your earlier observation of the spectrum of the coated end of the fluorescent lamp (or of any fluorescent lamp that is available). Describe the spectrum. Is it continuous? Is it the same as that of the incandescent lamp? How does it compare with the mercury spectrum?

________________________________________

You were able to identify the spectrum of the clear end of the lamp as that of mercury. Evidently, the mercury light somehow causes the coating to give off light that is very different from that of mercury itself.

Recall that the spectrum of the clear end of the *incandescent* lamp was not noticeably different from the spectrum of its coated end. Why does the coating of the fluorescent lamp behave very differently from the coating of the incandescent lamp? In order to answer this, you will need to learn about energy absorption and also what is meant by *fluorescence.*

### HOW ATOMS ACQUIRE ENERGY

You have now studied the relationship between the energy of an atom and the possible frequencies of the electromagnetic radiation which it can emit. You have also noted that atoms can absorb energy in fixed amounts. The atom is normally in the ground state, and thus the absorbed energy raises it to one of the excited states. If the atom is already in an excited state, then the absorbed energy raises it to a still higher excited state.

The *excitation process* is basically the same for atoms in a flame as it is for atoms in a fluorescent lamp or a discharge tube. In each of these light sources, atoms are excited by collisions with other atoms or with electrons.

In the flame, heat energy is shared by all the atoms, causing their kinetic energies to increase. As they move faster, any one atom has more frequent and more violent collisions with other atoms.

In some of these collisions, the atoms simply rebound from each other like billiard balls, with no change in the *internal* energy of either atom. In other collisions, the *internal* energy of one or both atoms increases by some allowed amount. For some cases, this energy increase may even be sufficient to ionize the atom; that is, to remove an electron from the atom. Since the remaining atom is positively charged (is a *positive ion*) and the electron is negatively charged (is a *negative ion*), we say that an *ion pair* is produced.

The excitation process in a fluorescent lamp is much the same. Electrons inside the lamp are accelerated by electric forces. (Electric energy is converted to the kinetic energy of the electrons.) As these electrons move about in the mercury vapor, they collide with mercury atoms, with any of the results mentioned. If the mercury atom becomes ionized, the electrical forces will also accelerate the newly freed electron and the positive mercury ion. These in turn gain energy and produce more collisions which produce more excited atoms and more ions. This series of processes is called a *gas discharge.* It has much in common with a spark.

**Question 19.** The discussion of the discharge process may have suggested that the build-up

of ion pairs would continue indefinitely. Why is this not possible?

**Question 20.** As the discharge process develops, why do the positive ions and negative electrons not recombine into neutral atoms?

There is another way by which atoms become excited. This second type of excitation process involves the absorption of energy as electromagnetic radiation. This process, which is essential to the understanding of the fluorescent lamp, is illustrated in the next experiment.

## A BRIEF STUDY OF ABSORPTION SPECTRA

### EXPERIMENT C-2. The Absorption Spectrum of Sodium

As you did before, light the bunsen burner and place several crystals of rock salt in the flame, using forceps. You have already measured the spectrum of the yellow light emitted by the sodium atoms in the flame. Place an incandescent light source behind the flame so that you can look *through the flame* toward the filament, being sure that the filament of the lamp is vertical (Figure 23).

Using your diffraction grating, look for a very thin black line in the yellow portion of the continuous spectrum of the lamp. While looking at the black line, observe the effects of removing the bunsen burner. Then, with the burner in place, observe the effects of turning off the light source. How does the "wavelength" of the black line compare with that of the bright yellow line in the flame spectrum?

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The spectrum you have just observed was generated by shining white light (with its continuous spectrum) through heated sodium vapor (with its bright line spectrum). The dark line represents a frequency which has been *removed* from the continuous spectrum. Apparently the sodium atoms can absorb only

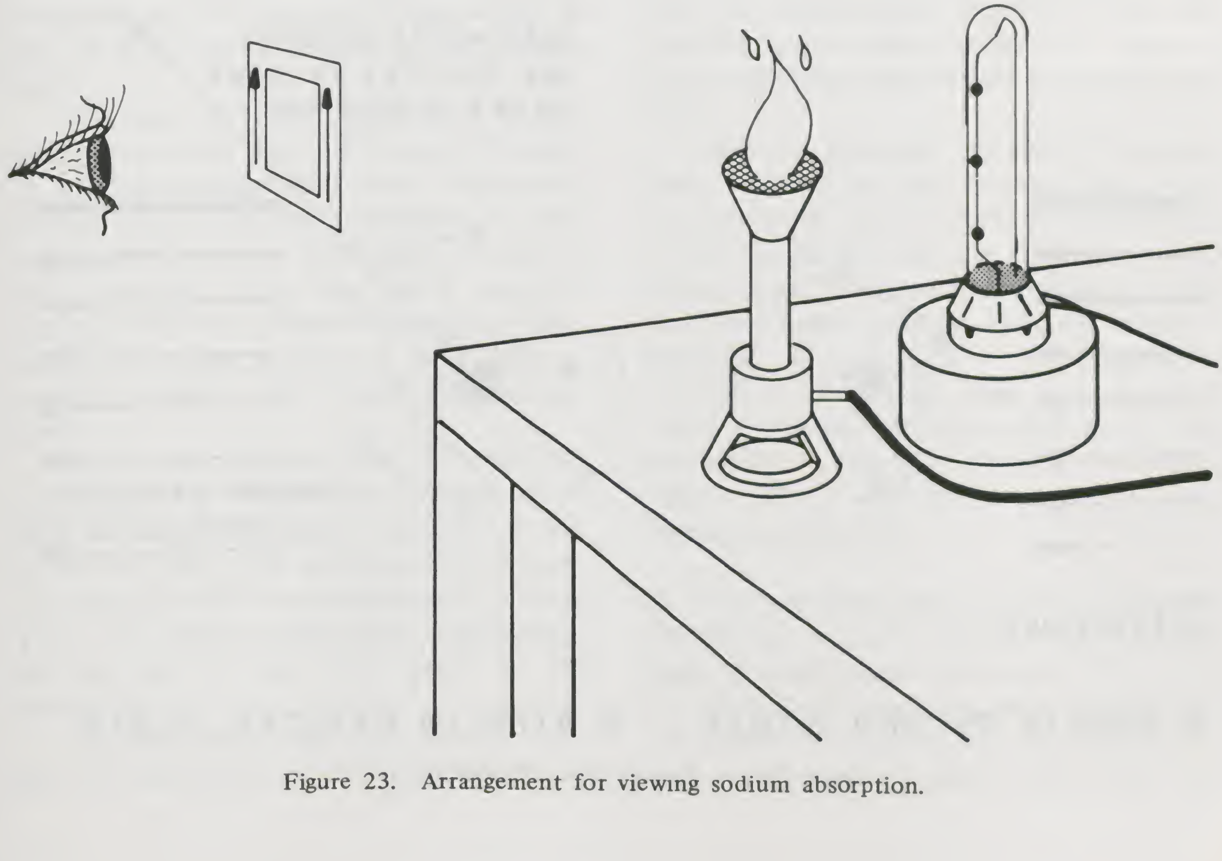

Figure 23. Arrangement for viewing sodium absorption.

certain frequencies from the continuous spectrum provided by the white light. The light which was absorbed by the sodium atoms is re-emitted, but in all directions. So the result is that the yellow light from the lamp, which was going in the forward direction toward the observer, is scattered in all directions and very little of it reaches the diffraction grating. This produces a dark line where the yellow used to appear. Such a dark line spectrum is called an *absorption spectrum.* Figure 24 represents this process.

For a sodium atom, the frequency of the absorbed radiation is the same as that of the visible light which is emitted in the flame. This frequency corresponds to a transition between the ground state and first excited state of the sodium atom. This is reasonable in view of the fact that atoms are normally in the ground state. (Even in the flame, most of the sodium atoms are in the ground state.)

Soon after an atom absorbs radiation (and thus is raised to an excited state), it returns to the ground state by emitting radiation. The frequency of the *emitted* radiation may be the same as that of the *absorbed* radiation. However, if there are intermediate excited states, the atom may "stop on the way." In that case the atom emits two (or more) frequencies whose sum is equal to that of the absorbed radiation.

Absorption spectra are important in the understanding of many atomic phenomena. For example, when the light emitted by the very hot gases in the core of the sun passes through the cooler gases on the surface of the sun, an absorption spectrum is produced. Each of the cooler gases can absorb particular frequencies from the continuous spectrum produced by the sun. The element helium (atomic number 2) was discovered by analysis of the sun's absorption spectrum before it was discovered on earth. (The name *helium* comes from *helios,* the Greek word for the sun.)

**Question 21.** How can spectral analysis lead to information on

a. what gases are contained in an automobile exhaust,
b. heavy elements in the sun?

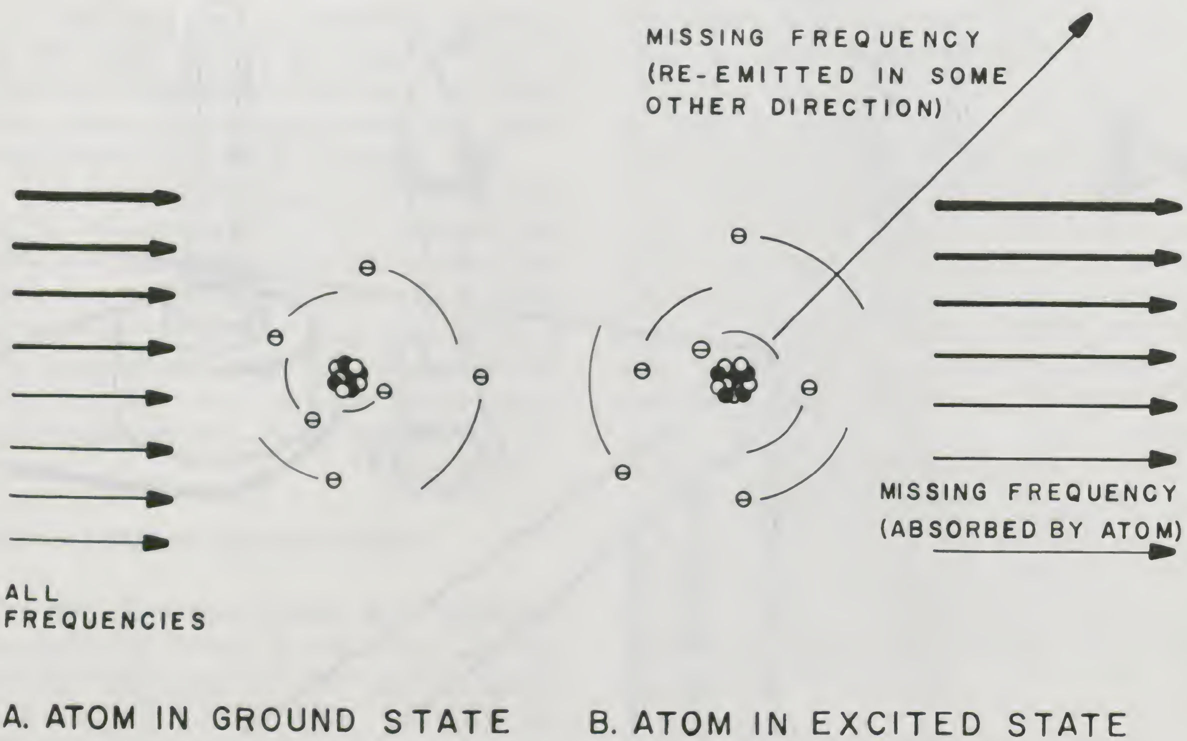


Figure 24. A representation of absorption.

**Question 22.** After an atom absorbs radiation it re-radiates it, often at the same frequency. Why does this re-radiation not "cancel" the absorption line?

## FLUORESCENCE

We shall now return to the fluorescent lamp and learn more about what takes place inside it.

The spectrum from the coated end of the fluorescent lamp is similar to the continuous spectrum of the incandescent lamp. When the lamp is turned off (mercury atoms are not emitting), the white coating emits no light. We have seen that atoms can absorb energy (either from atomic collisions or light energy) and thus they can be excited to higher energy states. We have also seen that the atoms can then undergo transitions to lower energy levels, emitting light in the process. The material on the inner surface of the glass tube of the fluorescent lamp must be able to absorb some of the energy emitted by the mercury vapor and re-emit that energy in the form of light which appears to be white. This process of the absorption of light at one frequency and the emitting of the energy as light of lower frequencies is called *fluorescence.*

The fact that the visible lines of the mercury spectrum can still be seen through the coating indicates that these frequencies are not being absorbed appreciably by the coating on the lamp. What kind of light is being absorbed? Since the visible mercury spectrum does not contain frequencies which excite the coating, it must be some kind of invisible radiation that causes the fluorescence.

When you photographed the mercury spectrum, it was pointed out that the *ultraviolet* radiation is the most important in the fluorescent lamp. The principal line of mercury is the one with a wavelength of 253.7 nm. (This is far shorter than visible wavelengths, and the line is not even visible in the photograph.)

Materials which have the ability to absorb one wavelength and then to emit light of longer wavelengths are called *phosphors.* The white coating inside the fluorescent tube is a particular phosphor which absorbs ultraviolet light and emits visible light.

## EXPERIMENT C-3. Phosphors and Black Lights

A. You will find in your apparatus two identical packages of chalk. Place one package under the clear end of the fluorescent lamp and the other under the clear end of the incandescent lamp. With both lamps on, darken the room and carefully examine the chalk. Are all of the pieces of chalk still white? What happens after the lamps are turned off (with the room dark)?

We have already stated that ultraviolet light is necessary to excite the fluorescent lamp phosphor. The chalk you used is made of other kinds of phosphors, which give off colored rather than white light when excited. Since the fluorescent lamp caused the chalk to glow more brightly (to *fluoresce* more strongly), it seems logical to assume that the fluorescent spectrum contains more of the radiation which excites these phosphors than does the incandescent lamp. A typical fluorescent lamp emits about five times as much ultraviolet light as does an incandescent lamp.

When a particular phosphor is excited, some of the atoms may remain excited for a long period of time. That is, all of the atoms do not immediately undergo transitions to the ground state. This results in emission of light for some period of time after the energy is absorbed. The property of a phosphor to continue to glow after it has been excited is called *persistence.* The persistence of the phosphor of a fluorescent lamp is very short. That is why it seems to go out immediately when it is turned off.

B. Ask your instructor for a *black light.* (Be careful not to look directly into the black light; it could injure your eyes.) Shine it on the chalk (with the room darkened). What is the effect? Your instructor will provide you with a variety of phosphors which will "turn

on" in the presence of the black light. Experiment with these in various ways. You might use the black light to excite various phosphors and try to observe their persistences. Another thing is to look at various white objects, such as paper, white shirts, etc., under the black light. Can you explain what happens?

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What is a black light? You cannot measure its spectrum with your grating spectrometer because the lamp is constructed so that visible light is filtered out. (This accounts for the name *black light.*) If you could measure the spectrum, you would find that the light emits much ultraviolet radiation, usually from mercury vapor, as in the fluorescent lamp. A typical black light is a fluorescent lamp whose tube is constructed of a material which transmits (rather than absorbs) the ultraviolet light. An external filter blocks out most of the visible light.

**Question 23**. You observed the sodium absorption spectrum. How is that process different from fluorescence?

**Question 24**. Clocks often have "luminous dials." Some glow dimly all night, but all of them glow brightly for several minutes after the lights are turned off. Explain the process involved during these few minutes.

**Quesiton 25**. Can you explain why certain white objects, such as bond paper and white shirts, fluoresce very strongly in the presence of a black light?

**Problem 6**. Referring to Figure 20, suppose radiation of wavelength 160.7 nm (equivalent to 7.73 eV) shines through mercury vapor. Some of this radiation is absorbed by the mercury vapor. What wavelengths might be radiated as a result of the absorption (You did some of the necessary calculations earlier.) Why might this be called the "fluorescence of mercury vapor"?

**Problem 7**. The electrons in a TV tube have energies of about 20,000 eV. These electrons are stopped by the atoms in the phosphor coating on the inside surface. If we suppose that all the energy of an electron is converted to "light," what would be its wavelength? Referring to Figure 7, decide what kind of radiation it would be.

### Spectra of Phosphors

Obviously, phosphors are available which emit light in many colors, ranging from the white of the fluorescent lamp through the various pastel colors of the chalk, to the vivid colors of fluorescent poster paints. You have observed the apparently continuous spectrum of the phosphor in the fluorescent lamp.

What about the spectra of the other phosphors? These are very difficult to measure since the amount of light radiated is so small, but you would find them rather different from the line spectra of gases. Line spectra are characteristic of atoms in a gas where the only atomic interactions are collisions. In solids or liquids (or gases in which two or more atoms are bound together into molecules), the atoms interact with each other in a more complicated way, causing the spectral lines to broaden into *spectral bands.* It should be clear that this will be true for both emission and absorption processes. These effects can be more simply demonstrated by absorption experiments, such as the following.

### EXPERIMENT C-4. Band Spectra

You are provided with several test tubes which contain various solutions. Let white light shine through each test tube and look at its absorption spectrum with your diffraction grating.

Describe your observations. Can you in any way relate the color of the solution to the absorption spectrum you see? You will also be given a discharge tube containing carbon dioxide ($CO_2$) and a tube containing iodine ($I_2$), both of which are molecular, not

atomic, gases. Observe their emission spectra with your grating. Do they contain only sharp lines?

As discussed earlier, the line spectra of gases are caused by transitions of electrons in the region of a single atomic nucleus. If two or more atoms combine to form a molecule, each electron is then under the influence of two or more nuclei. Also, each nucleus is under the influence of the other nuclei in the molecule. These effects greatly complicate the energy-level diagram of a molecule, compared with that of an atom. The highly complex energy structure is responsible for the complicated spectra of molecular materials. *Many* transitions become possible, and the spectral lines resulting from these transitions may be close enough together to form bands rather than lines. In a liquid or solid phosphor, the energy structures are even more complicated. So, although the spectrum of a molecular gas may consist of some lines and some bands, the spectra of solids and liquids almost always consists of bands. In the case of the phosphor in the fluorescent lamp, these emission bands are close enough together to form what resembles a continuous spectrum.

## PRACTICAL FLUORESCENT LAMPS

Let's now put all of these ideas together and look at a real fluorescent lamp. Figure 25 identifies most of the important components of a practical fluorescent lamp.

The electrodes (both called *cathodes* in this case) are made of a material which emits electrons easily. These electrons are accelerated by the applied voltage and initiate the discharge process described earlier. As more ions are created, they are accelerated by the voltage between the cathodes. This causes the formation of still more ions. This process would continue without limit, causing larger and larger amounts of charge to flow. To prevent this high build-up, a current-limiting device, called a *ballast*, is included in the external circuit. Usually this is just a large coil of wire which prevents excessively large currents.

You have probably noticed a "hesitation" in the operation of many fluorescent lamps. This time lag between the closing of the switch and the actual lighting of the lamp can be understood by looking at the fluorescent lamp which you have been using.

## EXPERIMENT C-5. The Fluorescent Lamp

Turn on the fluorescent lamp. Carefully observe the electrodes when you hold down the "start" switch for a few seconds. Does any obvious thing happen before the lamp will trigger? What is necessary for the lamp to turn on?

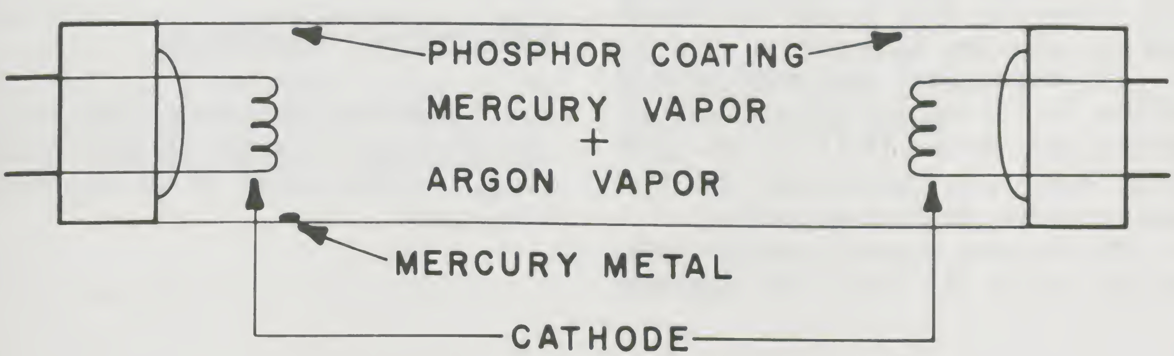


Figure 25. Details of a fluorescent lamp.

It should be clear from this observation that the electrodes are heated until they become slightly incandescent. The bluish glow around each electrode is a small mercury discharge (called a *glow discharge*) caused by the rapid emission of electrons from the heated cathodes.

When you release the switch, the ballast creates a very short duration, high voltage pulse. This voltage is large enough to accelerate the electrons from the vicinity of the heated cathode toward the opposite end of the lamp, causing the entire lamp to *discharge* as previously described. Once the discharge is started, it will continue to operate on ordinary line voltage.

In practical fluorescent lamps, a *starter switch* is often used to provide both preheating of the cathodes and the high voltage pulse. The delay between closing the wall switch and seeing the lamp light is due to this *starter.* It is simply an automatic version of the *start* switch you have been using. It should be pointed out that there are many fluorescent lamps for which no perceptible delay occurs between the closing of the switch and the lighting of the lamp. Such *rapid-start* lamps depend upon a variety of circuit designs to supply the necessary initial starting voltage and the initial glow discharge around the cathodes.

You may have also noticed the flickering of a fluorescent lamp as the lamp fails. Usually, the electrodes are still intact, but they are no longer able to emit enough electrons to maintain the current through the lamp. The starter repeatedly tries to start the lamp causing the lamp to turn on and off until the starter itself burns out.

The lamp contains some argon vapor in addition to the mercury vapor. There are several reasons for this. One is that the argon ionizes more easily than mercury and thus assists in starting the discharge process.

The phosphor coating is carefully selected to convert the 253.7 nm ultraviolet radiation from the mercury vapor into light which is very nearly "white". If you compare commercially available fluorescent lamps, you will find several different phosphors in use. The light from some of these lamps is "cool" to your eyes. That is, the blue end of the spectrum is strongest. Others are "warm," which implies an emphasis on the red end of the spectrum.

In terms of what is actually radiated, how does a fluorescent lamp compare with an incandescent lamp? The graphs in Figure 26 illustrate the difference for typical lamps.

With the same power input, the fluorescent lamp emits about four times as much visible light as an incandescent lamp and less than one half as much infrared radiation. As previously mentioned, infrared radiation is essentially heat radiation. This means that incandescent lamps will be more effective in heating a room than fluorescent lamps.

In Experiment A-1a, you probably observed that the fluorescent lamp became hotter to the touch than the incandescent lamp. Yet it radiated less heat. You should now be able to understand this observation. The ultraviolet light absorbed by the glass envelope is changed into heat and the glass becomes hot. For the incandescent lamp, the filament emits a large amount of infrared radiation, which the glass does not stop, but which does heat your hand when it is a short distance from the lamp.

**Question 26.** Why does the phosphor have to be on the *inside* surface of the fluorescent lamp bulb?

**EXPERIMENT C-6. (Optional)**

Remove the bottom cover from the lamp box and draw a diagram of the electrical circuit, showing details of the fluorescent lamp circuit.

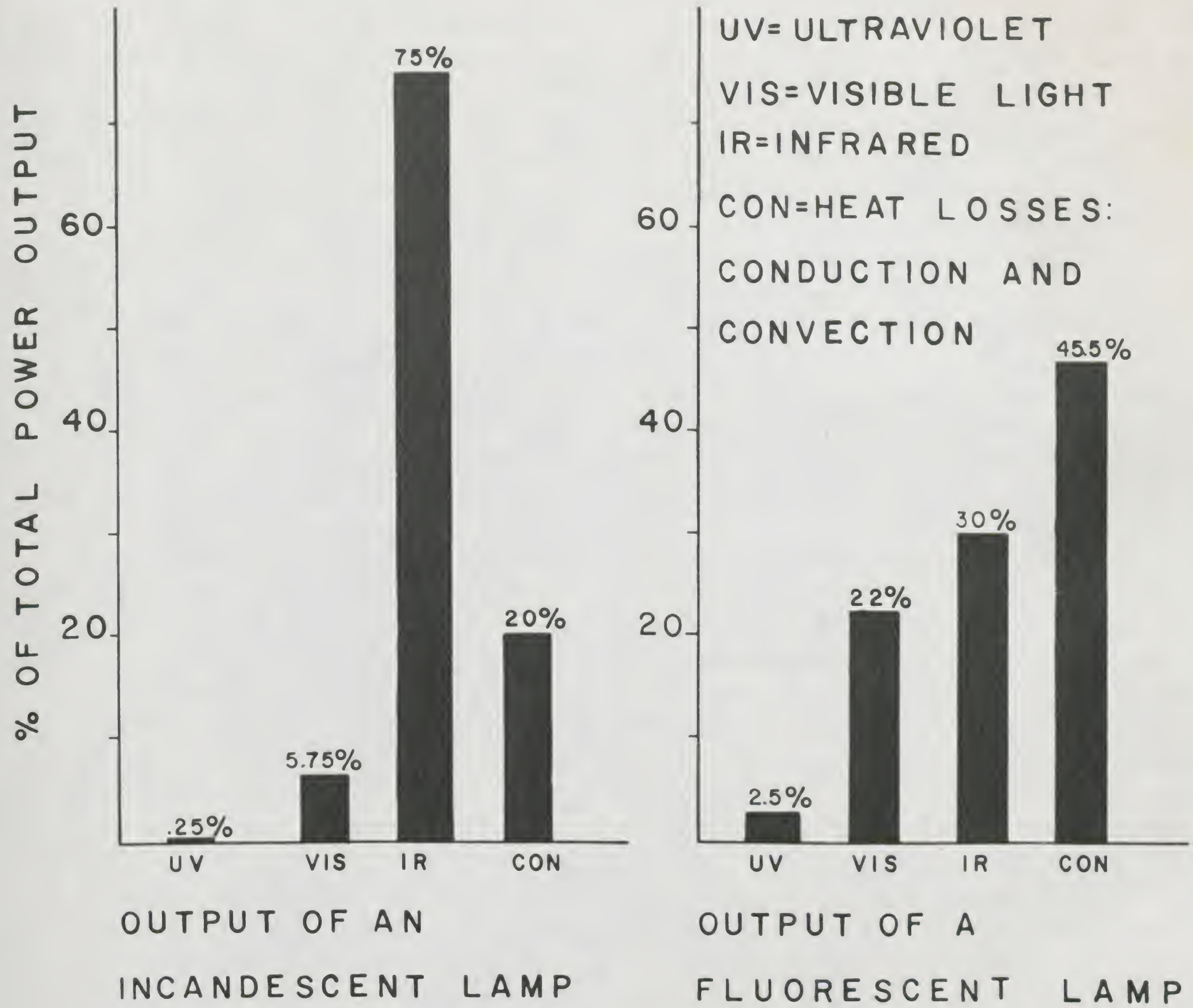


Figure 26. Radiation from fluorescent and incandescent lamps.

## SUMMARY

In this final section of the module you studied energy absorption processes. You saw how an intense ultraviolet line in the mercury spectrum can account for the fluorescence of the coating on the inside surface of the fluorescent lamp. The processes of absorption and re-emission of light were also seen to be important in the fluorescence of many objects when illuminated by *black light.*

Finally, the construction of commercial fluorescent lamps was discussed and the functions of the various parts were pointed out. Fluorescent lamps were then compared with incandescent lamps, in terms of the various kinds of radiations emitted.

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